PERIODIC TABLE

1.0 INTRODUCTION :

The arrangement of all the known elements according to their properties in such a way that the elements with similar properties are grouped together in a tabular form is called periodic table.

DEVELOPMENT OF PERIODIC TABLE

(A) LAVOISIER CLASSIFICATION :

(i) Lavoisier classified the elements simply in metals and non metals.

Metals are the one which have the tendency of losing the electrons.

$Na \rightarrow Na^+ + e^-$ and $K \rightarrow K^+ + e^-$

Non-metals are the one which have the tendency of gaining the electrons.

$F + e^- \rightarrow F^-$ and $Cl + e^- \rightarrow Cl^-$

(ii) **Drawback or Limitation :**

(a) As the number of elements increases, this classification became insufficient for the study of elements.

(b) There are few elements which have the properties of both metals as well as non-metals and they are called metalloids. Lavoisier could not decide where to place the metalloids.

(B) PROUT'S HYPOTHESIS :

He simply assumed that all the elements are made up of hydrogen, so we can say that

Atomic weight of element = n × (Atomic weight of one hydrogen atom)

$$\boxed{\text{Atomic weight of H} = 1}$$

where n = number of hydrogen atom = 1, 2, 3,....

Drawback or Limitation :

(i) Every element can not be formed by Hydrogen.

(ii) Atomic weight of all elements were not found as the whole numbers.

Ex. Chlorine (atomic weight 35.5) and Strontium (atomic weight 87.6)

(C) DOBEREINER TRIAD RULE [1817] :

(i) He made groups of three elements having similar chemical properties called TRIAD.

(ii) In Dobereiner triad, atomic weight of middle element is nearly equal to the average atomic weight of first and third element.

Ex.

Cl	Br	I	Ca	Sr	Ba	Li	Na	K
35.5	80.0	127	40	87.6	137	7	23	39

$$\left[\; x = \frac{35.5+127}{2} = 81.2 \right. \qquad \left[\; x = \frac{40+137}{2} = 88.5 \right. \qquad \left[\; x = \frac{7+39}{2} = 23 \right.$$

Where x=average atomic weight

(iii) Other examples – (K, Rb, Cs), (P, As, Sb), (S, Se, Te)

Drawback or Limitation : All the known elements could not be arranged as triads. It is not applicable for d and f-block elements.

(D) NEWLAND OCTAVE RULE [1865]

(i) He arranged the elements in the increasing order of their atomic mass and observed that the properties of every 8^{th} element was similar to the 1^{st} element. (like in the case of musical vowels notation)

(ii) At that time inert gases were not known.

Sa	Re	Ga	Ma	Pa	Dha	Ni	Sa
1	2	3	4	5	6	7	8

						H
Li	Be	B	C	N	O	F
Na	Mg	Al	Si	P	S	Cl
K	Ca					

(iii) The properties of Li are similar to 8^{th} element i.e. Na and Be are similar to Mg and so on.

Drawback or Limitation :

(a) This rule is valid only upto Ca because after Ca due to presence of d-block element there is a difference of 18 elements instead of 8 elements.

(b) After the discovery of Inert gas and including them in the periodic table, it has become the 8^{th} element from Alkali metal so this law had to be dropped out.

(E) LOTHAR MEYER'S CURVE [1869] :

(i) He plotted a curve between atomic weight and atomic volume of different elements.

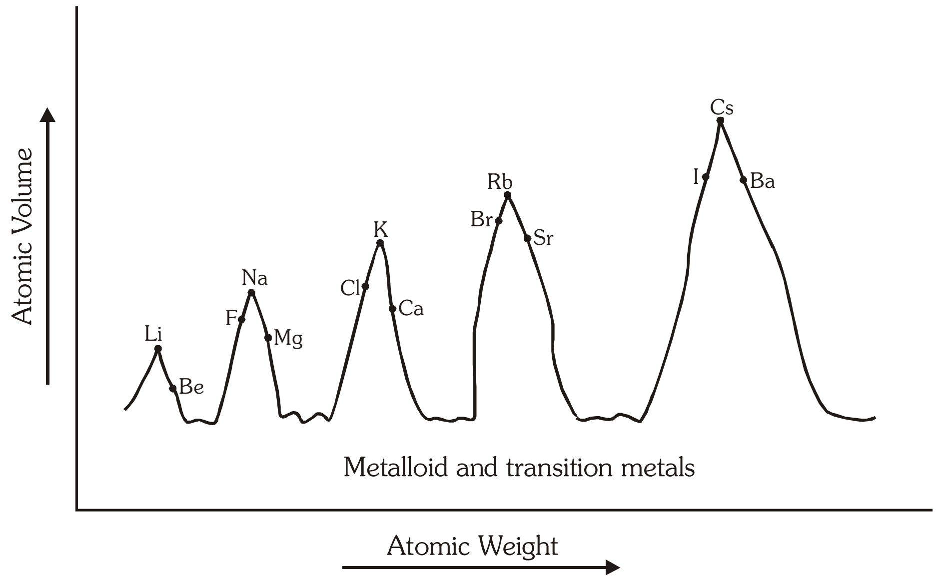
(ii) The following observation can be made from the curve –

- (a) Most electropositive elements i.e. alkali metals (Li, Na, K, Rb, Cs) occupy the peak positions on the curve.
- (b) Less electropositive i.e. alkaline earth metal (Be, Mg, Ca, Sr, Ba) occupy the descending position on the curve.
- (c) Metalloids (Si, Ge, As, Sb, Te, Po, At) and transition metals occupy bottom part of the curve.
- (d) Most electronegative i.e. halogens (F, Cl, Br, I) occupy the ascending position on the curve.

Note : Elements having similar properties occupy similar position on the curve.

Conclusion : On the basis of this curve Lother Meyer proposed that the physical properties of the elements are periodic function of their atomic weight and this has become the base of Mendeleev's periodic table.

Periodic function : When the elements are arranged in the increasing order of their atomic weight, elements having similar properties gets repeated after a regular interval.



(F) MENDELEEV'S PERIODIC TABLE [1869] :

(i) Mendeleev's periodic law : The physical and chemical properties of elements are the periodic function of their atomic weight.

(ii) Characteristics of Mendeleev's periodic table :

(a) It is based on atomic weight

(b) 63 elements were known, noble gases were not discovered.

(c) He was the first scientist to classify the elements in a systematic manner i.e. in horizontal rows and in vertical columns.

(d) Horizontal rows are called periods and there were 7 periods in Mendeleev's Periodic table.

(e) Vertical columns are called groups and there were 8 group in Mendeleev's Periodic table.

(f) Each group upto VII is divided into A & B subgroups.'A' sub group element are called normal or representative elements and 'B' sub group elements are called transition elements.

(g) The VIII group consisted of 9 elements in three rows (Transitional metals group).

(h) The elements belonging to same group exhibit similar properties.

(iii) Merits or advantages of Mendeleev's periodic table :

(a) Study of elements : First time all known elements were classified in groups according to their similar properties. So study of the properties of elements become easier .

(b) Prediction of new elements : It gave encouragement to the discovery of new elements as some gaps were left in it.

Sc (Scandium) Ga (Gallium) Ge (Germanium) Tc (Technetium)

These were the elements for whom position and properties were well defined by Mendeleev even before their discoveries and he left the blank spaces for them in his table.

Ex. Blank space at atomic weight 72 in silicon group was called Eka silicon (means properties like silicon) and element discovered later was named Germanium .

Similarly other elements discovered after mendeleev's periodic table were.

Eka Aluminium	– Galium(Ga)	Eka Boron	– Scandium (Sc)
Eka Silicon	– Germanium (Ge)	Eka Mangenese	– Technetium (Tc)

(c) Correction of doubtful atomic weights : Correction were done in atomic weight of some elements.

$$\boxed{\text{Atomic weight} = \text{Valency} \times \text{Equivalent weight.}}$$

Initially, it was found that equivalent weight of Be is 4.5 and it is trivalent (V = 3), so the weight of Be was 13.5 and there is no space in Mendeleev's table for this element. So, after correction, it was found that Be is actually bivalent (V = 2). So, the weight of Be became 2 ×4.5 = 9 and there was a space between Li and B for this element in Mendeleev's table.

Corrections were done in atomic weight of elements are – **U, Be, In, Au, Pt.**

(iv) Demerits of Mendeleev's periodic table :

(a) Position of hydrogen : Hydrogen resembles both, the alkali metals (IA) and the halogens (VIIA) in properties so Mendeleev could not decide where to place it.

(b) Position of isotopes : As atomic wt. of isotopes differs, they should have placed in different position in Mendeleev's periodic table. But there were no such places for isotopes in Mendeleev's table.

(c) Anomalous pairs of elements : There were some pair of elements which did not follow the increasing order of atomic weights.

Ex. Ar and Co were placed before K and Ni respectively in the periodic table, but having higher atomic weights.

$\begin{pmatrix} \text{Ar} & \text{K} \\ 39.9 & 39.1 \end{pmatrix}$ $\begin{pmatrix} \text{Te} & \text{I} \\ 127.5 & 127 \end{pmatrix}$

$\begin{pmatrix} \text{Co} & \text{Ni} \\ 58.9 & 58.6 \end{pmatrix}$ $\begin{pmatrix} \text{Th} & \text{Pa} \\ 232 & 231 \end{pmatrix}$

(d) Like elements were placed in different groups :

There were some elements like Platinum (Pt) and Gold (Au) which have similar properties but were placed in different groups in Mendeleev's table.

Pt	Au
VIII	IB

(e) Unlike elements were placed in same group :

I[St] group

IA IB

Li

Na (Bridge Element)

More reactive	K	Cu	Less reactive
Alkali metal	Rb	Ag	Coinage metal
Normal elements	Cs	Au	Transition element

Cu, Ag and Au placed in I[st] group along with Na, K etc. While they differ in their properties (Only similar in having ns^1 electronic configuration)

(f) It was not clear that **'lanthanides** and **Actinides'** are related with IIIA group or IIIB group.

(g) Cause of periodicity : Why physical & chemical properties repeated in a group.

BEGINNER'S BOX–1

1. Mendeleev's periodic law is based on -

(1) Atomic number (2) Atomic weight (3) Number of neutrons (4) None of the above

2. The first attempt to classify elements systematically was made by -

(1) Mendeleev (2) Newland (3) Lother Meyer (4) Dobereiner

3. Atomic weight of an element X is 39, and that of element Z is 132. atomic weight of their intermediate element Y, as per dobereiner triad, will be

(1) 88.5 (2) 93.0 (3) 171 (4) 85.5

4. Which of the following is not a dobereiner triad

(1) Li, Na, K (2) Mg, Ca, Sr (3) Cl, Br, I (4) S, Se, Te

5. The law of triads is applicable to

(1) C, N, O (2) H, O, N (3) Na, K, Rb (4) Cl, Br, I

6. The law of triads is not applicable on

(1) Cl, Br, I (2) Na, K, Rb (3) S, Se, Te (4) Ca, Sr, Ba

7. Which of the following set of elements obeys Newland's octave rule –

(1) Na, K, Rb (2) F, Cl, Br (3) Be, Mg, Ca (4) B, Al, Ga

8. For which of the pair Newland octave rule is not applicable –

(1) Li, Na (2) C, Si (3) Mg, Ca (4) Cl, Br

9. Which of the following element was present in Mendeleev's periodic table?

(1) Sc (2) Tc (3) Ge (4) None of these

10. Is Fe, Co, Ni are dobereiner triad ?

1.1 MODERN PERIODIC TABLE (MODIFIED MENDELEEV PERIODIC TABLE) :

(i) It was proposed by Moseley.

(ii) Modern periodic table is based on atomic number.

(iii) Moseley did an experiment in which he bombarded high speed electron on different metal surfaces and obtained X-rays.

He found out that $\sqrt{\nu} \propto Z$ where ν = frequency of X-rays, Z = atomic number.

From this experiment, Moseley concluded that the physical and chemical properties of the elements are periodic function of their atomic number. It means that when the elements are arranged in the increasing order of their atomic number elements having similar properties gets repeated after a regular interval. This is also known as 'Modern Periodic Law'.

(iv) **Modern periodic law :** The physical & chemical properties of elements are the periodic function of their atomic number.

(v) **Characteristics of modern periodic table :**

(a) 9 vertical columns called groups.

(b) I to VIII group + 0 group of inert gases.

(c) Inert gases were introduced in periodic table by Ramsay.

(d) 7 horizontal rows called periods.

LONG FORM / PRESENT FORM OF MODERN PERIODIC TABLE :

(It is also called as 'Bohr, Bury, Rang & Werner Periodic Table)

(i) It is based on the Bohr-Bury electronic configuration concept and atomic number.

(ii) This model is proposed by Rang & Werner

(iii) 7 periods and 18 groups

(iv) According to I. U. P. A. C. 18 vertical columns are named as 1^{st} to 18^{th} group.

(v) The co-relation between the groups in long form of periodic table and in modern form of periodic table are given below –

IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIII			IB	IIB	IIIA	IVA	VA	VIA	VIIA	0
1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18

(vi) Elements belonging to same group have same number of electrons in the outermost shell so their properties are similar.

Description of periods

Period	n	Period Sub shell	No. of elements	Element	Name of Period
1.	1	1s	2	$_{1}H - _{2}He$	Shortest
2.	2	2s, 2p	8	$_{3}Li - _{10}Ne$	Short
3.	3	3s, 3p	8	$_{11}Na - _{18}Ar$	Short
4.	4	4s, 3d, 4p	18	$_{19}K - _{36}Kr$	Long
5.	5	5s, 4d, 5p	18	$_{37}Rb - _{54}Xe$	Long
6.	6	6s, 4f, 5d, 6p	32	$_{55}Cs - _{86}Rn$	Longest
7.	7	7s, 5f, 6d, 7p	32	$_{87}Fr - _{118}Uuo$	Complete

CONCLUSION

1. Period number = outermost shell
2. Number of element in a period = Number of electrons in a period subshell

DESCRIPTION OF GROUPS :

1st/IA/Alkali metals

H = $1s^1$

Li = $1s^2$, $2s^1$

Na = $1s^2$, $2s^2\ 2p^6$, $3s^1$

K = $1s^2$, $2s^2\ 2p^6$, $3s^2\ 3p^6$, $4s^1$

General electronic configuration = ns^1

Number of valence shell e^- = 1

2nd/IIA/Alkaline earth metals

Be = $1s^2$, $2s^2$

Mg = $1s^2$, $2s^2\ 2p^6$, $3s^2$

Ca = $1s^2$, $2s^2\ 2p^6$, $3s^2\ 3p^6$, $4s^2$

General electronic configuration = ns^2

(n = Number of shell)

Number of valence shell e^- = 2

13th/IIIA/Boron Family

B = $1s^2$, $2s^2\ 2p^1$

Al = $1s^2$, $2s^2\ 2p^6$, $3s^2\ 3p^1$

Ga = $1s^2$, $2s^2\ 2p^6$, $3s^2\ 3p^6\ 3d^{10}$, $4s^2\ 4p^1$

General electronic configuration = $ns^2\ np^1$

Number of valence shell e^- = 3

14th/IVA/Carbon Family

C = $1s^2$, $2s^2\ 2p^2$

Si = $1s^2$, $2s^2\ 2p^6$, $3s^2\ 3p^2$

Ge = $1s^2$, $2s^2\ 2p^6$, $3s^2\ 3p^6\ 3d^{10}$, $4s^2\ 4p^2$

General electronic configuration = $ns^2\ np^2$

Number of valence e^- = 4

15th/VA/Nitrogen family/Pnicogen

(Used in fertilizer as urea)

N = $1s^2$, $2s^2\ 2p^3$

P = $1s^2$, $2s^2\ 2p^6$, $3s^2\ 3p^3$

As = $1s^2$, $2s^2\ 2p^6$, $3s^2\ 3p^6\ 3d^{10}$, $4s^2\ 4p^3$

General electronic configuration = $ns^2\ np^3$

Number of valence shell e^- = 5

16th/VIA/Oxygen family/Chalcogen

(Ore forming)

O = $1s^2$, $2s^2\ 2p^4$

S = $1s^2$, $2s^2\ 2p^6$, $3s^2\ 3p^4$

Se = $1s^2$, $2s^2\ 2p^6$, $3s^2\ 3p^6\ 3d^{10}$, $4s^2\ 4p^4$

General electronic configuration = $ns^2\ np^4$

Number of valence shell e^- = 6

17^{th}/VIIA/Fluorine family/Halogens	**18^{th}/Zero group/Inert gases / Noble gases**
(Salt forming) F = $1s^2, 2s^2\ 2p^5$ Cl = $1s^2, 2s^2\ 2p^6, 3s^2\ 3p^5$ Br = $1s^2, 2s^2\ 2p^6, 3s^2\ 3p^6\ 3d^{10}, 4s^2\ 4p^5$ General electronic configuration = $ns^2\ np^5$ Number of valence shell e^- = 7	(Less reactive) Ne = $1s^2, 2s^2\ 2p^6$ Kr = $1s^2, 2s^2\ 2p^6, 3s^2\ 3p^6\ 3d^{10}, 4s^2\ 4p^6$ General electronic configuration = $ns^2\ np^6$ (except He) Number of valence shell e^- = 8

(vii) NOMENCLATURE OF ELEMENTS :

(a) IUPAC gave names to elements above atomic number 100 as follows –

0	1	2	3	4	5	6	7	8	9
nil	un	bi	tri	quad	pent	hex	sept	oct	enn

(b) In all the elements suffix is – ium.

Ex.

Atomic No.	**IUPAC Name**	**Symbol**	**Elemental Name**	**Symbol**
101	Un nil unium	Unu	Mendelevium	Md
102	Un nil bium	Unb	Nobelium	No
103	Un nil trium	Unt	Lawrencium	Lr
104	Un nil quadium	Unq	Rutherfordium	Rf
105	Un nil pentium	Unp	Dubnium	Db
106	Un nil hexium	Unh	Seaborgium	Sg
107	Un nil septium	Uns	Bohrium	Bh
108	Un nil octium	Uno	Hassium	Hs
109	Un nil ennium	Une	Meitnerium	Mt
110	Un un nilium	Uun	Darmstadtium	Ds

(viii) Identification of group, period and block :

(A) When atomic number is given :

Step I : $71 \geq Z \geq 58 \Rightarrow$ Lanthanoids (6^{th} Period)] f-block

$103 \geq Z \geq 90 \Rightarrow$ Actinoids (7^{th} Period)

Group number = IIIB (largest group of periodic table)

Step II : Z = 104 to 118 (Period number = 7)

Group number = last two digits in atomic number of element

Example : Z = 104

Group no. = 4

Step III : **Group number** = 18 + given atomic number – atomic number of next noble gas

If the value of this formula is negative then use 32 instead of 18 in formula.

(B) When electronic configuration is given

Period number (n) = number of outermost shell/Highest shell number.

Block identification :

- If np electron present then p - block (ns^2 np^{1-6})

  group number = 12 + np electrons
- If np electron absent then s/f/d block

  If $(n-2)f^0$ $(n-1)d^0$ ns^{1-2} = s block

  group number = ns electrons

  If $(n-2)f^{1-14}$ $(n-1)d^{0-1}$ ns^2 = f block

  group number = IIIB
- If any other configuration or $(n-1)d^{1-10}$ ns^{0-2} (d-block)

  group number = (n–1)d electron + ns electron

Bohr's Classification

Inert gases	**Normal or representative elements**	**Transition element**	**Inner transition element**
outermost shell complete 6 element	outermost shell incomplete s & p block element except inert gas 38 element	n & n–1 shells incomplete either in atomic or ionic form all d block element except = IIB (Zn, Cd, Hg & Uub) 36 element	n, (n–1), (n–2) shells incomplete f-block elements 28 elements

SOME IMPORTANT POINTS :

(a) 2nd period elements (Li, Be, B) Shows **diagonal relationship** with 3rd period elements (Mg, Al, Si). Because of same ionic potential value they shows similarity in properties.

$$\textbf{(Ionic potential } (\phi) = \frac{\textbf{Charge on cation}}{\textbf{Radius of cation}}\textbf{)}$$

$$\begin{array}{cccccc} Li & & Be & & B & \\ & \searrow & & \searrow & & \searrow \\ Na & & Mg & & Al & & Si \end{array}$$

(b) 3rd period elements (Na, Mg, Al, Si, P, S, Cl) except inert gases are called **typical elements** because they represent the properties of other element of their respective group.

(c) **TRANSURANIC ELEMENTS** :

Elements having atomic number more than 92 are known as transuranic element.

All transuranic elements are radioactive & artificial.

$$\begin{array}{ccc} & & & \text{f-block} & \text{d-block} \\ Th & Pa & U & \left[\begin{array}{ll} Np — & Lr \\ 93 & 103 \end{array}\right] & \left[\begin{array}{ll} Unq — & Uub \\ 104 & 112 \end{array}\right] \\ 90 & 91 & 92 & & \end{array}$$

First man made element is Tc

First man made lanthanoid is Pm

All actinoids are radioactive but all lanthanoids are not artificial / man made (except Pm)

(d) The group containing most electro positive elements – GROUP IA.

(e) The group containing most electro negative elements – GROUP VIIA

(f) The group containing maximum number of gaseous elements–GROUP ZERO(18th)

(g) The group in which elements have generally ZERO valency – GROUP ZERO(18^{th})

(h) **In the periodic table**

Number of Gaseous elements – 11 (H, N, O, F, Cl + Noble gases)

Number of Liquid elements – 6 (Cs, Fr, Ga, Hg, Br, Uub)

Number of Liquid elements at room temp. – 2

Bromine is the only non-metal which exists in liquid form.

Number of Solid elements – 95 (if discovered elements are 112)

(i) 0/18 group have all the elements in gaseous form.

(j) 2nd period contains maximum number of gaseous elements. They are 4 **(N, O, F, Ne)**

(k) IIIB/3rd group is called longest group having 32 elements including 14 Lanthanides and 14 Actinides

Sc

Y

La................Lanthanides (14)

Ac................Actinides (14)

BEGINNER'S BOX-2

1. Which of the following is best general electronic configuration of normal element

(1) $ns^{1-2}\ np^{0-6}$ (2) $ns^{1-2}\ np^{1-5}$ (3) $ns^{1-2}\ np^{0-5}$ (4) $ns^{1-2}\ np^{1-6}$

2. Which of the following set of atomic numbers represents representative element

(1) 5, 13, 30, 53 (2) 11, 33, 58, 84 (3) 5, 17, 31, 54 (4) 9, 31, 53, 83

3. Which of the following electronic configuration does not belongs to same block as others :-

(1) [Xe] $4f^{14}\ 5d^{10}\ 6s^2$ (2) [Kr] $4d^{10}\ 5s^2$ (3) [Kr] $5s^2$ (4) [Ar] $3d^6\ 4s^2$

4. The electronic configuration of an element is $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$. What is the atomic number of next element of the same group which is recently discovered :-

(1) 20 (2) 119 (3) 111 (4) None

5. Which of the following electronic configurations in the outermost shell is characteristic of alkali metals

(1) $(n-1)\ s^2p^6\ ns^2p^1$ (2) $(n-1)\ s^2p^6d^{10}\ ns^1$ (3) $(n-1)\ s^2p^6\ ns^1$ (4) $ns^2np^6\ (n-1)d^{10}$

6. Which of the following elements belong to alkali metals ?

(1) $1s^2,\ 2s^2\ 2p^2$ (2) $1s^2,\ 2s^2\ 2p^6,\ 3s^2\ 3p^6\ 3d^{10},\ 4s^2\ 4p^6,\ 5s^1$

(3) $1s^2,\ 2s^2\ 2p^5$ (4) None of these

7. Elements whose atoms have three outermost shells incomplete are called –

(1) s-block (2) p-block (3) d-block (4) f-block

8. Which of the following statement is wrong :-

(1) All the actinides are synthetic (man made) elements

(2) In the Lanthanides last electron enters in 4f orbitals

(3) Np_{93} onwards are transuranic elements

(4) Lanthanum is d–block element

9. Which of the following statement is wrong :-

(1) Total no. of liquid elements in the periodic table.....Six

(2) First metal element in the periodic table is....Li

(3) All type of elements are present in 6^{th} period

(4) Iodine is a gaseous element.

10. An element which is recently discovered is placed in 7^{th} period and 10^{th} group. IUPAC name of the element will be :-

(1) Unnilseptium (2) Ununnilium (3) Ununbium (4) None

1.2 PERIODICITY

(A) In a period, the ultimate orbit remain same, but the number of e^- gradually increases.

In a group, the number of e^- in the ultimate orbit remains same, but the values of n increases.

(B) Causes of periodicity :

(a) The cause of periodicity in properties is due to the same outermost shell electronic configuration coming at regular intervals.

(b) In the periodic table, elements with similar properties occur at intervals of 2, 8, 8, 18, 18 and 32. These numbers are called as magic numbers.

SCREENING EFFECT (σ) AND EFFECTIVE NUCLEAR CHARGE (Z_{eff}) :

(i) Valence shell e^- suffer force of attraction due to nucleus and force of repulsion due to inner shell electrons.

(ii) The decrease in force of attraction on valence e^- due to inner shell e^- is called screening effect or shielding effect.(i.e. total repulsive force is called shielding effect.)

(iii) Due to screening effect valence shell e^- experiences less force of attraction exerted by nucleus.

i.e. total attraction force experieced by valence electrons represented by a number is Z_{eff}.

(iv) There is a reduction in nuclear charge due to screening effect. Reduced nuclear charge is called effective nuclear charge.

(v) If nuclear charge = Z, effective nuclear charge = Z_{eff} , σ (Sigma)= Screening constant or shielding constant.

So, $\boxed{Z_{eff} = (Z - \sigma)}$

- **Slater's rule to know screening constant (σ)**

(a) For single electron species $\sigma = 0$

(b) Screening effect (S.E.) for two e^- species 0.30

Ex. In He ($1s^2$)

Screening effect of one 1s e^-. where $\sigma = 0.30$

$\therefore$ Zeff = $Z - \sigma = 2 - 0.30 = 1.7$

(c) Screening effect of each ns and np (Outermost orbit) electrons is 0.35

(d) Screening effect of each (n – 1) penultimate orbit s, p, d electrons is 0.85

(e) Screening effect of each (n – 2) and below all the e^- present in s, p, d, f is 1.0

From top to bottom in a group Z_{eff} remain constant

Group	Element	**Li**	**Na**	**K**	**Rb**	**Cs**	**Fr**
	Z_{eff}	1.30	2.20	2.20	2.20	2.20	2.20
Period	Element	**Be**	**B**	**C**	**N**	**O**	**F**
	Z_{eff}	1.95	2.6	3.25	3.90	4.55	5.20

For same shell shielding effect has the order as s > p > d > f (due to penetration effect)

Z_{eff} for different ions of an element

$\boxed{Z_{eff} \propto \dfrac{\text{positive charge}}{\text{negative charge}}}$ (i) Z_{eff} for different ions of an element
(ii) Z_{eff} for isoelectronic species.

(i) Z_{eff} for different ions of an element

Ex. $N^+ > N > N^- = Z_{eff}$

(ii) Z_{eff} of isoelectronic species

Ex. $H^- < Li^+ < Be^{+2} < B^{+3}$ ($2e^-$ species)

$N^{-3} < O^{-2} < F^- < Na^+ < Mg^{+2}$ ($10e^-$ species)

1.3 ATOMIC RADIUS

The average distance of valence shell e^- from nucleus is called atomic radius. It is very difficult to measure the atomic radius because –

(i) The isolation of single atom is very difficult.

(ii) There is no well defined boundary for the atom. (The probability of finding the e^- is 0 only at infinity).

So, the more accurate definition of atomic radius is –

- Half the inter-nuclear distance(d) between two atoms in a homoatomic molecule is known as atomic radius.
- This inter-nuclear distance is also known as bond length. Inter-nuclear distance depends upon the type of bond by which two atoms combine.

Based on the chemical bonds, atomic radius is divided into four categories –

(A) Covalent radius **(B) Ionic radius** **(C) Metallic radius** **(D) van der Waal's radius**

(A) Covalent Radius

(SBCR –Single Bonded Covalent Radius)

(a) Covalent bonds are formed by overlapping of atomic orbitals.

(b) Internuclear distance is minimum in this case.

(c) Covalent radius is the half of the internuclear distance between two singly bonded homo atoms.

Ex. If internuclear distance of A–A(A_2) molecule is (d_{A-A}) and covalent radius is r_A then

$$d_{A-A} = r_A + r_A \quad \text{or} \quad 2r_A$$

$$r_A = \frac{d_{A-A}}{2}$$

Ex. In Cl_2 molecule, internuclear distance is 1.98A^0 so $r_{cl} = \frac{1.98}{2} = 0.99\,\text{Å}$

(B) Ionic Radius

(i) Cationic Radius

(a) When an neutral atom loses e^- it converts into cation (+ve charged ion)

(b) Cationic radius is always smaller than atomic radius **because** after loosing e^- number of e^- reduces, but number of protons remains same, due to this Z_{eff} increases, hence electrons are pulled towards nucleus and atomic radius decreases, moreover after loosing all the electrons from the outer most shell, penultimate shell becomes ultimate shell which is nearer to nucleus so size decreases.

(c) Size of cation $\propto \dfrac{1}{\text{Magnitude of the charge or } Z_{eff}}$

Ex. (i) $Fe > Fe^{+2} > Fe^{+3}$

(ii) $Pb^{+2} > Pb^{+4}$

(iii) $Mn > Mn^{+2} > Mn^{+3} > Mn^{+4} > Mn^{+5} > Mn^{+6} > Mn^{+7}$

(ii) Anionic Radius

(a) When a neutral atom gains e^- it converts into anion [Negative charge ion]

(b) Anionic radius is always greater than atomic radius **because** in an anion e^- are more than protons and inter electronic repulsion increases, which also increases screening effect. So effective nuclear charge reduces, so distance between e^- and nucleus increases and size of anion also increases.

Ex. Flourine (Z=9)

	F	F^-
Proton	9	9
e^-	9	10

so $\frac{Z}{e} = \frac{9}{9} = 1$ $\quad \frac{9}{10} = 0.9$ As Zeff of F^- is less than F so size of $F^- > F$

(c) Size of isoelectronic species :

- Those species having same number of e^- but different nuclear charge forms isoelectronic series.
- For isoelectronic species the atomic radius increases with decrease in effective nuclear charge

Species	**K^+**	**Ca^{+2}**	**S^{-2}**	**Cl^-**
Z	19	20	16	17
e	18	18	18	18
$\frac{Z}{e}$	$\frac{19}{18}$	$\frac{20}{18}$	$\frac{16}{18}$	$\frac{17}{18}$

Order of radius : $\underbrace{(S^{-2} > Cl^- > K^+ > Ca^{+2})}_{18\ e^-}$, $\underbrace{(N^{3-} > O^{2-} > F^- > Na^+ > Mg^{+2} > Al^{+3})}_{10\ e-}$

(C) Metallic/Crystal Radius

(a) Half of the inter nuclear distance between two adjacent metallic atoms in crystalline lattice structure.

(b) there is no overlapping of atomic orbitals, So **Metallic radius > Covalent radius**

(c) Metallic radius $\propto \frac{1}{\text{Metallic bond strength}}$

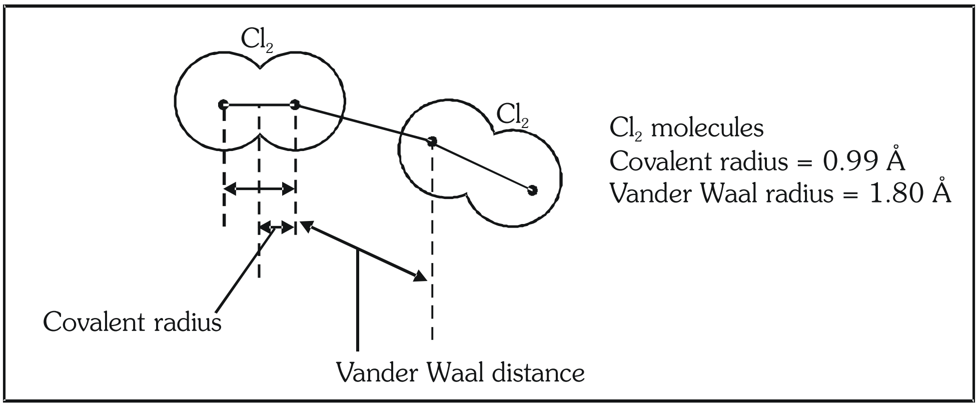
(D) Vander Waal's Radius

(a) Those atoms (like noble gases) which are not bonded with each other, experiences a weak attractive force to come nearer.

(b) Half of the distance between the nuclei of adjacently placed atoms in solid state of a noble gas is Vander Waal's radius.

(c) Inert gas have only Vander Waal radius.

(d) In molecules of nonmetals solid both covalent and van der Waal's radius exists.



$\boxed{\textbf{Vander Waal's radius} \cong \textbf{2} \times \textbf{ covalent radius}}$

$\boxed{\textbf{Vander Waal's radius > Metallic radius > Covalent radius}}$

- **Factors affecting atomic size are :**

  (a) **In a period**

  $$\text{Atomic radius} \propto \frac{1}{Z_{eff}} \propto \frac{\text{negative charge}}{\text{positive charge}}$$

  Li > Be > B > C > N > O > F

  (b) **In a group**

  Atomic radius $\propto$ number of shells

  Li < Na < K < Rb < Cs

- **Periodic variation of atomic size :**

  (i) **Across a period :** It decreases from left to right in a period as effective nuclear charge (Z_{eff})increases

  Ex. Li > Be > B > C > N > O > F

  (ii) **In a group :** It increases from top to bottom in a group as number of shell increases

  Ex. Li < Na < K < Rb < Cs

  Exceptions : Transition elements

  $$\underbrace{\text{Sc}\quad\text{Ti}\quad\text{V}\quad\text{Cr}\quad\text{Mn}}_{Z_{eff}\ >\ \text{Screening effect}} \qquad \underbrace{\text{Fe}\qquad\text{Co}\qquad\text{Ni}}_{Z_{eff}\ \approx\ \text{Screening effect}} \qquad \underbrace{\text{Cu}\qquad\qquad\text{Zn}}_{Z_{eff}\ <\ \text{Screening effect}}$$

  Order of covalent radius : Sc > Ti > V >Cr > Mn > Fe $\approx$ Co $\approx$ Ni < Cu < Zn

- **Lanthanide Contraction :**

  (a) Outermost electronic configuration of inner transition elements is

  $(n-2)\ f^{1-14}$, $(n-1)s^2p^6d^{0-1}$, ns^2 (n = 6 or 7)

  (b) e^- enters in (n – 2) f orbitals

  (c) Because of complicated structure of f-orbital and due to poor shielding f electrons, the outermost shell electrons get attracted towards nucleus.

  (d) In I[st], 2[nd] and 3[rd] transition series, Radii– 3d < 4d $\approx$ 5d **(except III[rd] B)**

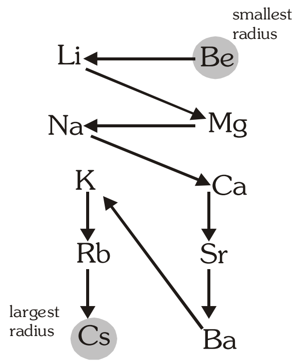
		IIIB	IVB	
size	↓	Sc	Ti ↓	size increases
increases		Y	Zr }	Almost equal due to lanthanide contraction
		La	Hf	

- **Transition contraction :**

  Note : While atomic size should increase down the group.

  At. size of Ga $\approx$ At. size of Al, due to transition contraction. (Due to poor shielding of d electrons)

- **s-block size variation**



BEGINNER'S BOX-3

1. From the given set of species, point out the species from each set having least atomic radius:-

(1) O^{-2}, F^-, Na^+ (2) Ni, Cu, Zn (3) Li, Be, Mg (4) He, Li^+, H^-

Correct answer is :-

(1) O^{-2}, Cu, Li, H^- (2) Na^+, Ni, Be, Li^+ (3) F^-, Zn, Mg, He (4) Na^+, Cu, Be, He

2. Which has the lowest anion to cation size ratio-

(1) LiF (2) NaF (3) CsI (4) CsF

3. Arrange the elements in increasing order of atomic radius Na, Rb, K, Mg :–

(1) Na, K, Mg, Rb (2) K, Na, Mg, Rb (3) Mg, Na, K, Rb (4) Rb, K, Mg, Na

4. Which of the following pairs of elements have almost similar atomic radii :-

(1) Zr, Hf (2) Mo, W (3) Co, Ni (4) All

5. If the ionic radii of $K^{\oplus}$ and $F^{\ominus}$ are nearly the same (i.e. 1.34 Å) then the atomic radii of K and F respectively are :–

(1) 1.34 Å, 1.34 Å (2) 0.72 Å, 1.96 Å (3) 1.96 Å, 0.72 Å (4) 1.96 Å, 1.34 Å

6. For the element X, student mansi measured its radius as 102 nm, student Rohit as 203nm. and Ankur as 100 nm. using same apparatus. Their teacher explained that measurements were correct by saying that recorded values by three students were :–

(1) Crystal, van der Waal and Covalent radii

(2) Covalent, crystal and van der Waal radii

(3) van der Waal, ionic and covalent radii

(4) None is correct.

7. Screening effect is not observed in :-

(1) He^+ (2) Li^{+2} (3) H (4) All of these

8. Arrange in orders of atomic and ionic radii :

(a) Ni , Cu , Zn (b) H^+ , H , H^- (c) Ti , Zr , Hf (d) Be , Li , Na

(e) Cr , V , Ti , Sc (f) I^+ , I , I^- (g) Sc , Y , La , Ac (h) Cl , Na , Rb

(i) Cu , Ag , Au (j) B , Be , Al , Mg (k) F , O , Cl , S

9. Which statement is false:-

(1) Screening effect increases down the group (2) Zeff increases down the group

(3) Zeff. increases in a period (4) All

10. The screening effect of d- electrons is :-

(1) Equal to the p - electrons (2) Much more than p - electrons

(3) Same as f - electrons (4) Less than p - electrons

1.4 IONISATION POTENTIAL OR IONISATION ENERGY OR IONISATION ENTHALPY (IP / IE)

(i) Minimum energy required to remove most loosely bonded outer most shell e^- in ground state from an isolated gaseous atom is known as ionization energy.

(Isolated $\rightarrow$ Without any bonding with other atom)

(ii) Successive Ionization Energy

(a) For an atom $M_{(g)}$ successive ionization energies are as follows -

$M_{(g)} + E_1 \longrightarrow M^+_{(g)} + e^-$ $\quad$ $E_1 = I^{st}$ Ionization energy

$M^+_{(g)} + E_2 \longrightarrow M^{+2}_{(g)} + e^-$ $\quad$ $E_2 = II^{nd}$ Ionization energy

$M^{+2}_{(g)} + E_3 \longrightarrow M^{+3}_{(g)} + e^-$ $\quad$ $E_3 = III^{rd}$ ionization energy

$\boxed{E_1 < E_2 < E_3 ....}$ (Always for an element)

(b) Electron can not be removed from solid state of an atom, it has to be convert into gaseous form, Energy required for conversion from solid state to gaseous state is called Sublimation energy.

(c) For any neutral atom ionization energy is always an endothermic process (ΔH = +ve)

(d) It is measured in eV/atom (electron volt/atom) or Kcal/mole or KJ/mole

FACTORS AFFECTING IONISATION ENERGY

In a period

(i) Effective nuclear charge (Z_{eff})

$$\boxed{\text{Ionisation Energy} \propto Z_{eff} \propto \frac{\text{positive charge}}{\text{negative charge}}}$$

Ion with high positive oxidation state will have high ionisation energy.

Ex. $Fe^{+3} > Fe^{+2} > Fe$

(ii) Stability of half filled and fully filled orbitals :

Half filled p^3,d^5, f^7 or fully filled p^6, d^{10}, f^{14} are more stable than others so it requires more energy.

IE_1 $\quad$ N $>$ O

$1s^2, 2s^2 2p^3$ $\quad$ $1s^2, 2s^2 2p^4$

After loosing one e^-, O attains electronic configuration of N, so II^{nd} ionisation energy of O is more than N.

IE_2 $\quad$ N $<$ O

$1s^2, 2s^2 2p^2$ $\quad$ $1s^2, 2s^2 2p^3$

(iii) Penetration power of sub shells

(a) Order of attraction of subshells towards nucleus (Penetration power) is $s > p > d > f$

(b) 's' subshell is more closer to nucleus so more energy will be required to remove e^- from s-subshell as comparison to p,d & f subshells.

Ex. **Be** $\quad$ **B**

$1s^2, 2s^2$ $\quad$ $1s^2, 2s^2 2p^1$

IE_1 $\quad$ Be $>$ B

After loosing one e^-, B attains electronic configuration of Be, so II^{nd} ionisation energy of B is more than Be.

IE_2 $\quad$ Be $<$ B

$1s^2, 2s^1$ $\quad$ $1s^2, 2s^2$

In a group

Atomic size : $IE \propto \dfrac{1}{\text{atomic size}}$

COMPARISON OF IONISATION ENERGY

(i) **In a period :** Z_{eff} increases so removal of electron becomes difficult and ionisation energy increases.

Order of IE of 2nd period elements $Li < B < Be < C < O < N < F < Ne$

(ii) **In a group :** Size increase so ionisation energy decrease.

$$\xrightarrow{\text{Li} \quad \text{Na} \quad \text{K} \quad \text{Rb} \quad \text{Cs}}$$

Size increases, Ionisation Energy decreases

Exception :

- Ionisation Energy Ga > Al (due to Transition contraction)
- Ionisation Energy of 5d > 4d(due to lanthanide contraction)

Ex. Hf > Zr

Application of ionisation energy :

(A) Metallic and non metallic character :

Generally for metals Ionisation Energy is low.

For Non-metals Ionisation Energy is high.

$$\boxed{\text{Metallic character} \propto \frac{1}{IE}}$$

(B) Reactivity of metals :

$$\boxed{\text{Reactivity of metals} \propto \frac{1}{IE}}$$

Reactivity of metals increases down the group as ionisation energy decreases.

(C) Stability of oxidation states of an element :

(a) If the difference between two successive ionisation energy of an element ≥ 16eV, then its lower oxidation state is stable.

Ex. $$\left.\begin{array}{l} Na_{(g)} \longrightarrow Na^{+}_{(g)} \\ Na^{+}_{(g)} \longrightarrow Na^{2+}_{(g)} \end{array}\right] \Delta IE = 42.7eV$$

Difference between ionisation energy > 16 eV. So Na^+ is more stable.

(b) If the difference between two successive ionisation energy of an element $\le$ 11 eV, then its higher oxidation state is stable.

Ex. $$\left.\begin{array}{l} Mg_{(g)} \longrightarrow Mg^{+}_{(g)} \\ Mg^{+}_{(g)} \longrightarrow Mg^{2+}_{(g)} \end{array}\right] \Delta IE = 7.4eV$$

Difference of ionisation energy < 11 eV. So Mg^{+2} is more stable.

Ex. $$\left.\begin{array}{l} Al_{(g)} \longrightarrow Al^{+}_{(g)} \\ Al^{+}_{(g)} \longrightarrow Al^{2+}_{(g)} \end{array}\right] \Delta IE = 12.8eV$$. So Al^+ is more stable

$$\left.\begin{array}{l} Al^{+}_{(g)} \longrightarrow Al^{2+}_{(g)} \\ Al^{2+}_{(g)} \longrightarrow Al^{3+}_{(g)} \end{array}\right] \Delta IE = 6.0\ eV$$. So Al^{+3} is more stable.

Overall order of stability is $\boxed{Al^{+3} > Al^{+} > Al^{+2}}$

(D) To determine the number of valence electron of an element :

Number of valence electrons = number of lower values of IP before 1st highest jump.

BEGINNER'S BOX-4

1. IP_1 and IP_2 of Mg are 178 and 348 K. cal mol^{-1}. The enthalpy required for the reaction $Mg \rightarrow Mg^{2+} + 2e^-$ is :-

(1) + 170 K.cal (2) + 526 K.cal (3) - 170 K.cal (4) - 526 K.cal

2. The IP_1, IP_2, IP_3, IP_4 and IP_5 of an element are 7.1, 14.3, 34.5, 46.8, 162.2 eV respectively. The element is likely to be:-

(1) Na (2) Si (3) F (4) Ca

3. Which of the following element has 2^{nd} IP < 1^{st} IP

(1) Mg (2) Ne (3) C (4) None

4. In which of the following the energy change corresponds to first ionisation potential only :-

(1) $X_{(g)} \rightarrow X^+_{(g)} + e$ (2) $X_{2(g)} \rightarrow X^+_{(g)} + e$

(3) $X_{(s)} \rightarrow X^+_{(g)} + e$ (4) $X_{(aq)} \rightarrow X^+_{(aq)} + e$

5. In the given process which oxidation state is more stable.

$$M_{(g)} \longrightarrow M_{(g)}^+ \quad IE_1 = 7.9 \text{ eV}$$

$$M_{(g)}^+ \longrightarrow M_{(g)}^{+2} \quad IE_2 = 15.5 \text{ eV}$$

(1) M^+ (2) M^{+2} (3) Both (4) None

6. The electronic configuration of some neutral atoms are given below :–

(A) $1s^2\ 2s^1$ (B) $1s^2\ 2s^2\ 2p^3$ (C) $1s^2\ 2s^2\ 2p^5$ (D) $1s^2\ 2s^2\ 2p^6\ 3s^1$

In which of these electronic configuration would you expect to have highest :–

(i) IE_1 (ii) IE_2

(1) C, A (2) B, A (3) C, B (4) B, D

7. The correct order of decreasing second ionization energy of Li, Be, Ne, C, B

(1) Ne>B>Li>C>Be (2) Li>Ne>C>B>Be

(3) Ne>C>B>Be>Li (4) Li>Ne>B>C>Be

8. In which of the following element has highest value of ionisation energy–

(1) Ti (2) Zr (3) Hf (4) None of these

9. What is the correct order of ionisation energy :

(1) $K < Cu < Cu^+ < K^+$ (2) $K < Cu^+ < Cu < K^+$

(3) $Cu^+ < K < Cu < K^+$ (4) $K^+ < Cu^+ < Cu < K$

10. Match the column.

	Column-I Valence electronic configuration	Column-II Successive ionisation energies
(a)	ns^1	(p) 19, 27, 36, 48, 270
(b)	ns^2	(q) 16, 28, 34, 260
(c)	$ns^2\ np^1$	(r) 18, 26, 230, 250
(d)	$ns^2\ np^2$	(s) 14, 200, 220, 240

1.5 ELECTRON AFFINITY/ELECTRON GAIN ENTHALPY (EA/ΔH_{eg})

(1) The amount of energy released when an electron is added to the valence shell of an isolated gaseous atom known as Electron affinity.

$$X_{(g)} + e^- \longrightarrow X^-_{(g)} + \text{Electron Affinity}$$

(2) Generally first electron addition of an isolated gaseous atom is an exothermic process (except stable electronic configuration)

$$X_{(g)} + e^- \longrightarrow X^-_{(g)} + \text{Electron Affinity}$$

ΔH_{eg_I} (first electron gain enthalpy) = – ve

(3) Second electron addition of an isolated gaseous atom is always an endothermic process due to inter electronic repulsion.

$$X^-_{(g)} + e^- \longrightarrow X^{2-}_{(g)}$$

$\Delta H_{eg_{II}}$ (second electron gain enthalpy) = positive

Ex.

$$O_{(g)} + e^- \longrightarrow O^-_{(g)} \; ; \quad \Delta H_{eg_I} = -141 \text{ KJ/mole}$$

$$O^-_{(g)} + e^- \longrightarrow O^{-2}_{(g)} \; ; \quad \Delta H_{eg_{II}} = +744 \text{ KJ/mole}$$

Net reaction $O_{(g)} + 2e^- \longrightarrow O^{-2}_{(g)}$ $\quad \Delta H_{eg_{II}} = +603$ KJ/mole

(4) Formation of poly negative anion like O^{-2}, N^{-3}, C^{-4} etc. is always an endothermic process.

(5) Electron affinity of neutral atom is equal to ionisation energy of its anion.

$$X_{(g)} + e^- \underset{\text{Ionisation Potential}}{\overset{\text{Electron Affinity}}{\rightleftarrows}} X^-_{(g)} \quad \left(\text{EA of } X_{(g)} = \text{IE of } X^-_{(g)}\right)$$

(6) IE of neutral atom is equal to electron affinity of its cation

$$X_{(g)} \underset{\text{Electronic affinity}}{\overset{\text{Ionisation potential}}{\rightleftarrows}} X^+_{(g)} \quad \left(\text{IE of } X_{(g)} = \text{EA of } X^+_{(g)}\right)$$

(7) **Factors affecting electron affinity :**

(A) **Atomic size :** Electron Affinity $\propto \dfrac{1}{\text{Atomic size}}$

(B) **Effective nuclear charge (Z_{eff}) :** Electron Affinity $\propto Z_{eff} \propto \dfrac{\text{positive charge}}{\text{negative charge}}$

(C) **Stability of completely filled or half filled orbitals** : Electron affinity of elements having full-filled or half filled configuration is very less or zero so for these elements electron gain enthalpy $\left(\Delta H_{(eg)}\right)$ will be positive.

Ex. (i) EA $\Rightarrow ns^1 > ns^2$ (ii) EA $\Rightarrow np^2 > np^3$

(8) **Variation of electron affinity :**

(i) **In 2nd period –**

Li $\boxed{\text{Be}}$ B C $\boxed{\text{N}}$ O F $\boxed{\text{Ne}}$

$$\underbrace{Ne < Be < N}_{\text{endothermic}} < B < Li < C < O < F$$

(ii) **In Group :**

Electron affinity of 3rd period element is greater than electron affinity of 2nd period elements of the respective group.

F — $[He]\ 2s^2 2p^5$

Cl — $[Ne]\ 3s^2 3p^5$

Due to small size of fluorine, **electron density around the nucleus increases**. The incoming electron suffers more repulsion. In case of chlorine electron density decreases due to large size, decreasing order of electron affinity

Cl > F > Br > I	S > O > P > N	Si > C > P > N

Note : N & P have low electron affinity due to stable half filled configuration.

BEGINNER'S BOX-5

1. The correct order of electron affinity is :-

(1) Be < B < C < N (2) Be < N < B < C (3) N < Be < C < B (4) N < C < B < Be

2. In the formation of a chloride ion, from an isolated gaseous chlorine atom, 3.8 eV energy is released, which would be equal to :-

(1) Electron affinity of Cl^- (2) Ionisation potential of Cl

(3) Electronegativity of Cl (4) Ionisation potential of Cl^-

3. $O_{(g)} + 2e^- \rightarrow O^{2-}{}_{(g)}$ $\Delta Heg = 603$ KJ/mole. The positive value of ΔHeg is due to :-

(1) Energy is released to add on 1 e^- to O^{-1} (2) Energy is required to add on 1 e^- to O^{-1}

(3) Energy is needed to add on $1e^-$ to O (4) None of the above is correct

4. The electron affinity values for the halogens shows the following trend :–

(1) F < Cl > Br > I (2) F < Cl < Br < I (3) F > Cl > Br > I (4) F < Cl > Br < I

5. The process requiring the absorption of energy is.

(1) $F \rightarrow F^-$ (2) $Cl \rightarrow Cl^-$ (3) $O \rightarrow O^{2-}$ (4) $H \rightarrow H^-$

6. Second electron affinity of an element is :–

(1) Always exothermic (2) Endothermic for few elements

(3) Exothermic for few elements (4) Always endothermic

7. Process, $Na^+_{(g)} \xrightarrow{I} Na_{(g)} \xrightarrow{II} Na_{(s)}$

(1) In (I) energy released, (II) energy absorbed (2) In both (I) and (II) energy is absorbed

(3) In both (I) and (II) energy is released (4) In (I) energy absorbed, (II) energy released

8. Which of the following configuration will have least electron affinity.

(1) ns^2np^5 (2) ns^2np^2 (3) ns^2np^3 (4) ns^2np^4

9. Which of the following will have the most negative electron gain enthalpy and which the least negative ?

(1) F, Cl (2) Cl, F (3) Cl, S (4) Cl, P

10. Which arrangement represents the correct order of electron gain enthalpy (with negative sign) of the given atomic species ?

(1) S < O < Cl < F (2) O < S < F < Cl (3) Cl < F < S < O (4) F < Cl < O < S

1.6 ELECTRONEGATIVITY (EN)

(i) The tendency of a covalently bonded atom to attract shared pair of electrons towards itself is called electronegativity.

(ii) A polar covalent bond of A – B may be broken as

A —— $B \longrightarrow A^{\delta-}$ —— $B^{\delta+}$ (Electronegativity A > Electronegativity B)

depending on their tendency to attract bonded electron.

(iii) **Difference between electronegativity and Electron Affinity :**

Electronegativity	Electron Affinity
• Tendency of an atom in a molecule to attract the bonded electrons • It is not an energetic term • It regularly increases in a period because not depend on stable electronic configuration • It has no unit	• Energy released when an electron is added to neutral isolated gaseous atom • It is an energetic term • It does not increases regularly in a period because depend on stable electronic configuration • It is measured in eV/atom or KJ mol^{-1} or K cal mole^{-1}

(iv) EN was explained by Pauling for the first time

Electronegativity of some other elements are as follows –

						H 2.1
Li 1.0	Be 1.5	B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.2	Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
K 0.8						Br 2.8
Rb 0.8						I 2.5
Cs 0.7						
Fr 0.7						

In Pauling's scale, elements having almost same electronegativity are-

N = Cl = 3.0
C = S = I = 2.5
P = H = 2.1
Be= Al = 1.5
K = Rb = 0.8
Cs= Fr = 0.7

Note : Small atoms are normally having more electronegativity than larger atoms.

(v) FACTORS AFFECTING ELECTRONEGATIVITY :

(A) **Atomic size**

$$\boxed{\text{Electronegativity} \propto \frac{1}{\text{Atomic size}}}$$

Ex. F > Cl > Br > I

(B) **Effective nuclear charge (Z_{eff})**

$$\boxed{\text{Electronegativity} \propto Z_{eff} \propto \frac{\text{positive charge}}{\text{negative charge}}}$$

Ex. $Mn^{+2} < Mn^{+4} < Mn^{+7}$
$O^{-2} < O^{-1} < O < O^{+1} < O^{+2}$
$Fe < Fe^{+2} < Fe^{+3}$
-->
$Z_{eff} \uparrow$ EN$\uparrow$

(C) **% s - character**

$$\boxed{\textbf{Electronegativity} \propto \textbf{\% s - Character}}$$

(vi) PERIODIC TABLE & ELECTRONEGATIVITY :

(A) Electronegativity decreases down the group.

(B) In period on moving from left to right electronegativity increases.

(C) Electronegativity of Cs and Fr are equal, it is because from $_{55}Cs$ to $_{87}Fr$ only one shell increases but nuclear charge (No. of proton) increases by +32, so effect of nuclear charge balanced the effect of increase in number of shell.

Electronegativity of F > Cl but Electron affinity of Cl > F

(D) In IIIA group, value of electronegativity is irregular when going down the group, because of transition contraction

Electronegativity of Ga > Electronegativity of Al

(vii) APPLICATION OF ELECTRONEGATIVITY :

(A) **Metallic and non metallic nature :**

Generally metals have low electronegativity and non metals have high electronegativity, so we can say metallic character increases down the group but decreases along a period.

$\boxed{\textbf{Non Metallic Nature} \propto \textbf{EN}}$

(B) **Bond energy** : By increasing difference in electronegativity of bonded atoms, bond length decreases and hence bond energy increases

$\boxed{\text{Bond energy} \propto \text{Electronegativity difference}}$

HF > HCl > HBr > HI

(C) **Schoemaker and Stevenson law**

If in a diatomic molecule electronegativities of A – B have more difference. Then actual bond length will be reduced. As per schoemaker & Stevenson– The reduction in bond length depends on the difference in electronegativities of atoms by following manner -

$d_{A-B} = r_A + r_B - 0.09\,(X_A - X_B)$

Here X_A is E.N. of A & X_B is E.N. of B

Ex. If bond length of F_2 = 1.44 Å, Bond length of H_2 = 0.74Å. Find out the bond length of H – F ? (EN of F is 4.0, EN of H is 2.1)

Solution. $d_{H-F} = r_F + r_H - 0.09\,(X_F - X_H)$

$\because r_F = 1.44 / 2 = 0.72$ Å, $r_H = 0.74/2 = 0.37$ Å

$\therefore d_{H-F} = 0.72 + 0.37 - 0.09\,(4.0 - 2.1)$

$= 1.09 - (0.09 \times 1.9) = 1.09 - 0.171 = 0.919$ Å

(D) **Acidic & Basic Strength :**

(i) Nature of hydrides :

$\boxed{\text{Stability of molecule} \propto \text{Bond energy}}$

Order of stability of hydrohalides :	HF > HCl > HBr > HI
Order of acidic strength :	HF < HCl < HBr < HI
In VA group :	$NH_3 < PH_3 < AsH_3 < SbH_3 < BiH_3$

$\longrightarrow$

Thermal stability decreases　　Acidic character increases

(ii) Nature of hydroxides :

$$A{-}O{-}H \begin{cases} \longrightarrow AO^- + H^+ \text{ [Acidic]} \\ \longrightarrow A^+ + OH^- \text{ [Basic]} \end{cases}$$

(a) As per Gallis,

(i) In AOH if electronegativity of A is more than 1.7 (Non metal) then it is acidic in nature.

(ii) If electronegativity of 'A' is less than 1.7 (metal) then AOH will be basic in nature

(b) If $X_A - X_0 \geq X_0 - X_H$ (X_A = EN of A) then AO bond will be more polar and will break up as

$A — OH \longrightarrow A^+ + OH^-$ It shows basic nature

Ex. In NaOH

$X_O - X_{Na}$ (2.6) > $X_O - X_H$ (1.4) So hydroxide is basic

(c) If $X_A - X_O \leq X_O - X_H$ then OH bond will be more polar and will break up as

$A — O — H \longrightarrow H^+ + AO^-$ It shows Acidic nature

In ClOH

$X_O - X_{Cl}$ (0.5) < $X_O - X_H$ (1.4) So hydroxide is acidic

(iii) Nature of oxides :

(a) Along a period acidic nature increases.

(b) Down the group basic nature increases

Li	Be	B	C	N	O	F
Na	Mg	Al	Si	P	S	Cl

$\longleftarrow\!\longrightarrow$

Basic Amphoteric Acidic

ie. when in periodic table the distance between the element and oxygen increases, basic character increases.

$NO_2 > ZnO > K_2O$

$\longrightarrow$

acidic character decreases

Note: $BeO, Al_2O_3, ZnO, SnO, PbO, SnO_2, PbO_2, Sb_2O_3$ etc. are amphoteric oxides.

CO, H_2O, NO, N_2O etc. are neutral oxides.

$\boxed{\text{Acidic strength of oxide and oxyacid} \propto \text{Electronegativity}}$

B_2O_3 CO_2 N_2O_5

$\longrightarrow$

EN increase, acidic nature increase.

$HNO_3 > H_2CO_3 > H_3BO_3$

$H_3PO_4 > H_3AsO_4 > H_3SbO_4$

$HOF > HOCl > HOBr > HOI$

$HClO_4 > HBrO_4 > HIO_4$

$N_2O_5 > P_2O_5 > As_2O_5$

$N_2O_3 > P_2O_3 > As_2O_3 > Sb_2O_3$

$SO_3 > SeO_3 > TeO_3$

$H_2SO_3 > H_2SeO_3 > H_2TeO_3$

$\boxed{\text{Acidic nature} \propto \text{oxidation state}}$

Acidic properties increases with increasing oxidation state of an element

$HClO_4 > HClO_3 > HClO_2 > HClO$

$H_2SO_4 > H_2SO_3$

$SO_3 > SO_2$

$HNO_3 > HNO_2$

$N_2O_5 > N_2O_3$

$Sb_2O_5 > Sb_2O_3$

(E) Nature of bonds :

(a) According to Hanny & Smith formula

% ionic character $= 16 (X_A - X_B) + 3.5 (X_A - X_B)^2$

Here X_A = Electronegativity of A

X_B = Electronegativity of B

If $X_A - X_B \geq 2.1$ Ionic % $> 50\%$ i.e. Ionic bond

If $X_A - X_B \leq 2.1$ Ionic % $< 50\%$ i.e. covalent bond

(b) According to Gallis

$X_A - X_B \geq 1.7$ Ionic

$X_A - X_B \leq 1.7$ Covalent

If $X_A = X_B$; then A – B will be non polar. **Ex.** H—H, F—F

If $X_A > X_B$ and difference of electronegativities is small then

$A^{\delta-}$ —— $B^{\delta+}$ bond will be polar covalent

Ex. H_2O ($H^{\delta+}$ —— $O^{\delta-}$ —— $H^{\delta+}$)

If $X_A >> X_B$ and $X_A - X_B$ difference of electronegativities is high then

A^- —— B^+ bond will be polar or ionic

(F) Nomenclature of inorganic compounds :

Prefix —— less electronegative element

Suffix —— More electronegative element

Ex. Cl_2O (Right) OCl_2 (Wrong)

In Dichloroxide the electronegativity of Cl is less than 'O' i.e. why Cl is in prefix position.

OF_2 Oxygen difluoride

ICl Iodine chloride

(G) Bond polarity $\propto \Delta EN$

Ex. HF > HCl > HBr > HI

(viii) ELECTRONEGATIVITY SCALE :

Mulliken scale : According to Mulliken electronegativity is average value of ionisation potential and electron affinity of an element,

$$\boxed{X_m = \frac{\text{Ionisation Potential} + \text{Electron Affinity}}{2}}$$

where X_p is electronegativity on the basis of Pauling scale.

- If ionisation potential and electron affinity are given in eV, then electronegativity by Mulliken on Pauling scale will be

$$\boxed{X_p = \frac{\text{Ionisation Potential} + \text{Electron Affinity}}{5.6}}$$

$$\boxed{X_p = \frac{X_m}{2.8}}$$

- If ionisation potential and electron affinity are given in K.cal/mole then

$$\boxed{X_p = \frac{\text{Ionisation Potential} + \text{Electron Affinity}}{2 \times 62.5}}$$

BEGINNER'S BOX-6

1. Which of the following is affected by stable configuration of an atom :-

(a) Electronegativity (b) Ionisation potential (c) Electron affinity

Correct answer is :–

(1) Only electronegativity

(2) Only ionisation potential

(3) Electron affinity and ionisation potential

(4) All of the above

2. Which of the following elements have the different value of electronegativity :-

(1) H (2) S (3) Te (4) P

3. Which is the correct order of electronegativity –

(1) $Cl > S > P > Si$ (2) $Si > Al > Mg > Na$

(3) $F > Cl > Br > I$ (4) All

4. Electronegativity scale of pauling is based upon :-

(1) Bond length (2) Bond energy (3) Atomic radius (4) Covalent radius

5. Correct order of electronegativity of N, P, C and Si is :–

(1) $N < P < C < Si$ (2) $N > C > Si > P$

(3) $N = P > C = Si$ (4) $N > C > P > Si$

6. Outermost electronic configuration of the most electronegative element is :–

(1) ns^2np^3 (2) ns^2np^6 (3) ns^2 (4) ns^2np^5

7. Electronegativity of the following elements increases in the order.

(1) O, N, S, P (2) P, S, N, O (3) P, N, S, O (4) S, P, N, O

8. Give the correct order of electronegativity of central atom in following compounds –

(a)CH_3-CH_3, (b)$CH_2=CH_2$ (c) $CH\equiv CH$

The correct order is –

(1) a > b > c (2) c > a > b (3) c > b > a (4) b > c > a

ANSWER KEY

BEGINNER'S BOX-1	Que.	1	2	3	4	5	6	7	8	9	10
	Ans.	**2**	**1**	**4**	**2**	**4**	**2**	**3**	**4**	**4**	**No**

BEGINNER'S BOX-2	Que.	1	2	3	4	5	6	7	8	9	10
	Ans.	**3**	**4**	**3**	**3**	**3**	**2**	**4**	**1**	**4**	**2**

BEGINNER'S BOX-3	Que.	1	2	3	4	5	6	7	8	9	10
	Ans.	**2**	**4**	**3**	**4**	**3**	**1**	**4**		**2**	**4**

BEGINNER'S BOX-4	Que.	1	2	3	4	5	6	7	8	9	10
	Ans.	**2**	**2**	**4**	**1**	**2**	**1**	**4**	**3**	**1**	

BEGINNER'S BOX-5	Que.	1	2	3	4	5	6	7	8	9	10
	Ans.	**2**	**4**	**2**	**1**	**3**	**4**	**3**	**3**	**4**	**2**

BEGINNER'S BOX-6	Que.	1	2	3	4	5	6	7	8		
	Ans.	**3**	**2**	**4**	**2**	**4**	**4**	**2**	**3**		

EXERCISE-I (Conceptual Questions)

DEVELOPMENT OF PERIODIC TABLE

1. Mendeleev's periodic table is based on :-
(1) Atomic number
(2) Increasing order of number of protons
(3) Electronic configuration
(4) None of the above

2. Which of the following is/are Dobereiners triad :-
(a) P, As, Sb (b) Cu, Ag, Au
(c) Fe, Co, Ni (d) S, Se, Te
Correct answer is :-
(1) a and b (2) b and c (3) a and d (4) All

3. Which of the following sets of elements follows Newland's octave rule :-
(1) Be, Mg, Ca (2) Na, K, Rb
(3) F, Cl, Br (4) B, Al, Ga

4. Which are correct match :-
(a) Eka silicon – Be
(b) Eka aluminium – Ga
(c) Eka mangenese – Tc
(d) Eka scandium – B
(1) b, c (2) a, b, d (3) a, d (4) All

5. Atomic wt. of P is 31 and Sb is 120. What will be the atomic wt. of As, as per Dobereiners triad rule :-
(1) 151 (2) 75.5
(3) 89.5 (4) Unpredictable

6. The places that were left empty by Mendeleev's were, for:-
(1) Aluminium & Silicon
(2) Galium and germinium
(3) Arsenic and antimony
(4) Molybdenum and tungstun

7. Which is not anomalous pair of elements in the Medeleev's periodic table:-
(1) Ar and K (2) Co and Ni
(3) Te and I (4) Al and Si

8. The law of triads is applicable to :-
(1) Os, Ir, Pt (2) Ca, Sr, Ba
(3) Fe, Co, Ni (4) Ru, Rh, Pt

9. Elements which occupied position in the lother meyer curve, on the peaks, were :–
(1) Alkali metals
(2) Highly electro positive elements
(3) Elements having large atomic volume
(4) All

10. In a period the elements are arranged in :–
(1) Decreasing order of nuclear charge
(2) Decreasing order of No. of electrons
(3) Increasing order of nuclear charge
(4) In order of same nuclear charge

11. Which of the following statement is wrong :-
(1) 2^{nd} period contain 8 elements
(2) 3^{rd} period contains 18 elements
(3) 1^{st} period contains two non metals
(4) In p–block, metal, nonmetal and metalloids are present

12. Which of the following element was absent in the Mendeleev's periodic table :-
(1) Tc (2) Si
(3) B (4) F

13. IUPAC name of the element placed just after actinide series :-
(1) Unniltrium (2) Unnilpentium
(3) Unnilquadium (4) Ununbium

14. Which statement is wrong for the long form of periodic table :-
(1) Number of periods are 7 and groups 18
(2) No. of valence shell electrons in a period are same
(3) III^{rd} B group contains 32 elements
(4) Lanthanides and actinides are placed in same group

15. The elements which are cited as an example to proove the validity of mendeleev's periodic law are
(1) H, He (2) Ga, Sc
(3) Co, Ni (4) Zr, Hf

16. Which pair of successive elements follows increasing order of atomic weight in mendeleev's periodic table
(1) Argon and potassium
(2) Lithium and Beryilium
(3) Cobalt and nickel
(4) Tellurium and iodine

17. Which of the following statement is false :-
(1) Elements of ns^2np^6 electronic configuration lies in 1^{st} to 6^{th} period
(2) Typical elements lies in 3^{rd} period
(3) The seventh period will accommodate thirty two elements
(4) Boron and silicon are diagonally related

18. Among the Lanthanides the one obtained by synthetic method is :–
(1) Lu (2) Pm (3) Pr (4) Ce

PERIOD, GROUP AND BLOCK

19. Which of the following set of elements belongs to same period :-
(1) Zn, Cd, Hg (2) Fr, Ra, U
(3) K, Ca, Ag (4) None

20. The element with atomic number Z = 115 will be placed in :-
(1) 7^{th} period, IA group (2) 8^{th} period, IVA group
(3) 7^{th} period, VA group (4) 6^{th} period, VB group

21. Elements upto atomic no. 112 have been discovered till now. What will be the electronic configuration of the element possessing atomic no. 108 :-
(1) $[Rn]5f^{14}\ 6d^6\ 7s^2$ (2) $6f^{14}\ 7d^8\ 7s^2$
(3) $[Rn]\ 5f^{14}\ 6d^8\ 7s^0$ (4) $[Xe]\ 4f^{14}\ 5d^8\ 6s^2$

22. In 6th period of the modern periodic table, electronic energy levels are in the order :-
(1) 6s, 4f, 5d, 6p (2) 6s, 6p, 4f, 5d
(3) 4f, 5d, 6s, 6p (4) None

23. Out of first 100 elements no. of elements having electrons in 3d orbital (in their complete electronic configuration) are :-
(1) 80 (2) 100 (3) 40 (4) 60

24. The IUPAC name of the element which is placed after Db_{105} is the periodic table, will be :-
(1) Un nil pentium (2) Un un nilium
(3) Un nil hexium (4) Un nil quadium

25. The element with the electronic configuration $ns^2(n-1)s^2p^6d^0(n-2)s^2p^6d^{10}f^7$ lies in the :-
(1) s - block (2) p - block
(3) d - block (4) f - block

26. The element with atomic number Z=118 will be :–
(1) Noble gas
(2) Transition metal
(3) Alkali metal
(4) Alkaline earth metal

27. The atom having the valence shell electronic configuration $4s^2\ 4p^2$ would be in:-
(1) Group II A and period 3
(2) Group II B and period 4
(3) Group IV A and period 4
(4) Group IV A and period 3

28. The electronic configuration of d-block elements is exhibited by :-
(1) $ns^{1-2}(n-1)d^{1-10}$ (2) $ns^2\ (n-1)\ d^{10}$
(3) $(n-1)d^{10}s^2$ (4) ns^2np^5

29. The electronic configuration of the element with atomic number 109 if discovered will be:-
(1) $(n–1)d^7ns^2$ (2) $(n–1)d^9ns^2$
(3) nd^7ns^2 (4) $(n–1)d^5ns^2np^2$

30. The element having electronic configuration $4f^{14}\ 5d^0\ 6s^2$ belongs to :-
(1) d–block, 12^{th} group
(2) f–block, III B group
(3) f–block, 14^{th} group
(4) s–block, 2^{nd} group

31. Element with the electronic configuration given below, belong to which group in the periodic table $1s^2$, $2s^22p^6$, $3s^23p^63d^{10}$, $4s^24p^64d^{10}$, $5s^25p^3$
(1) 3^{rd} (2) 5^{th}
(3) 15^{th} (4) 17^{th}

32. $4d^35s^2$ configuration belongs to which group :-
(1) IIA (2) IIB (3) V B (4) III B

33. Which of the following electronic configuration belongs to inert gas elements :-
(1) $ns^2\ (n-1)d^{10}$ (2) $ns^2\ (n-1)s^2p^6$
(3) $ns^2\ np^6$ (4) None

34. From atomic number 58 to 71, elements are placed in ;-
(1) 5^{th} period and III A group
(2) 6^{th} period and III B group
(3) separate period and group
(4) 7^{th} period and IV B group

35. True statement is :-
(1) All the transuranic elements are synthetic elements
(2) Elements of third group are called bridge elements
(3) Element of $1s^2$ configuration is placed in IIA group
(4) Electronic configuration of elements of a group is same

36. Elements having $ns^2\ np^6$ valence shell electronic configuration lies in :-
(1) '0' gp. and $1^{st}–7^{th}$ period
(2) 18^{th} gp. and $2^{nd}–6^{th}$ period
(3) 18^{th} gp and $1^{st}–6^{th}$ period
(4) All are correct

37. Which of the following match is correct :-
(1) Last natural element – Uub
(2) General electronic configuration of IA group –ns^2
(3) Inert gas elements lies in 2^{nd} – 6^{th} period
(4) Typical elements – 3^{rd} period elements

38. The electronic configuration of elements X and Z are $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$ and $1s^2\ 2s^2\ 2p^5$ respectively. What is the position of element X with respect to position of Z in the periodic table -
(1) Just below Z (2) Just above Z
(3) Left to the Z (4) right to the Z

39. Which of the following sequence contains atomic number of only representative elements
(1) 55, 12, 18, 53 (2) 13, 33, 54, 83
(3) 3, 33, 53, 87 (4) 22, 33, 55, 66

40. Uranium (At No. - 92) is the last natural element in the periodic table. The last element of the periodic table which is recently discovered is Uub. What will be the total number of transuranic elements in the periodic table :-
(1) 21 (2) 20 (3) 11 (4) 12

41. Which two elements are in same period as well as same group of modern periodic table :-
(1) Z = 23, Z = 31 (2) Z = 65, Z = 66
(3) Z = 52, Z = 87 (4) Z = 58, Z = 46

42. Which of the following statement is not correct for given electronic configuration
$1s^2, 2s^2 2p^6, 3s^2 3p^6 3d^{10}, 4s^2 4p^6 4d^{10} 4f^{14}, 5s^2 5p^6\ 5d^{10}, 6s^2$
(1) It belongs to IIB group and 6^{th} period
(2) It is liquid at room temperature
(3) It is a transition element
(4) It is not used in high temperature thermometer

43. General electronic configuration of outermost and penultimate shell is $(n-1)\ s^2\ (n-1)p^6\ (n-1)\ d^x\ ns^2$. If n = 4 and x = 5, then number of protons in the nucleus will be :-
(1) > 25 (2) < 24 (3) 25 (4) 30

44. An ion M^{+3} has electronic configuration [Ar] $3d^{10}\ 4s^2$ element M belongs to :-
(1) s-block (2) p-block
(3) d-block (4) f-block

45. What is the atomic number of element having maximum no. of unpaired e^- in 4p subshell :-
(1) 33 (2) 17
(3) 53 (4)15

Zeff, SCREENING CONSTANT & ATOMIC ADIUS

46. The formula for effective nuclear charge is (if σ is screening constant)
(1) $Z - \sigma$ (2) $Z + \sigma$ (3) $Z\sigma^{-1}$ (4) $Z\sigma$

47. According to Slater rule, Effective nuclear charge in group generally :–
(1) Increases down the group
(2) Decreases down the group
(3) Remains constant
(4) First increases then decreases

48. In sodium atom the screening is due to :-
(1) $3s^2, 3p^6$ (2) $2s^1$
(3) $1s^2, 2s^2, 2p^6$ (4) $1s^2, 2s^2$

49. If the difference in atomic size of :
Na – Li = x; Rb – K = y; Fr – Cs = z
Then correct order will be:-
(1) x = y = z (2) x > y > z
(3) x < y < z (4) x < y << z

50. The correct order of size would be:-
(1) Ni < Pd ≃ Pt (2) Pd < Pt < Ni
(3) Pt > Ni > Pd (4) Pd > Pt > Ni

51. Which of the following order of radii is correct
(1) Li < Be < Mg (2) $H^+ < Li^+ < H^-$
(3) O < F < Ne (4) $Na^+ > F^- > O^{-2}$

52. K^+, Ar, Ca^{2+} and S^{2-} contains -
(1) Same electronic configuration and atomic volume
(2) Different electronic configuration but same IP.
(3) Same electronic configuration but different atomic volume
(4) None

53. Which of the following is not isoelectronic series :-
(1) Cl^-, P^{3-}, Ar (2) N^{3-}, Ne, Mg^{+2}
(3) B^{+3}, He, Li^+ (4) N^{3-}, S^{2-}, Cl^-

54. Which group of atoms have nearly same atomic radius:-
(1) Na, K, Rb, Cs (2) Li, Be, B, C
(3) Fe, Co, Ni (4) F, Cl, Br, I

55. Atomic radii of Fluorine and Neon in Angstrom units are given by :-
(1) 0.72, 1.60 (2) 1.60, 1.60
(3) 0.72, 0.72 (4) None of these

56. Which of the following has largest radius :-
(1) $1s^2\ 2s^2\ 2p^6\ 3s^2$
(2) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$
(3) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$
(4) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$

57. Which of the following order of atomic/ionic radius is not correct :–
(1) $I^- > I > I^+$ (2) $Mg^{+2} > Na^+ > F^-$
(3) $P^{+5} < P^{+3}$ (4) $Li > Be > B$

58. In the lithium atom screening effect of valence shell electron is caused by-
(1) Electrons of K and L shell
(2) Electrons of K shell
(3) Two electrons of 1^{st} and one of 2^{nd} shell
(4) None

59. Correct order of ionic radii is
(1) $Ti^{4+} < Mn^{7+}$ (2) $^{37}Cl^- < {}^{35}Cl^-$
(3) $K^+ > Cl^-$ (4) $P^{3+} > P^{5+}$

60. The radius of potassium atom is 0.203 nm. The radius of the potassium ion in nanometer will be :–
(1) 0.133 (2) 0.231
(3) 0.234 (4) 0.251

61. S^{-2} is not isoelectronic with :-
(1) Ar (2) Cl^-
(3) HS^- (4) Ti^{+3}

62. The best reason to account for the general tendency of atomic diameters to decrease as the atomic numbers increase within a period of the periodic table is the fact that
(1) Outer electrons repel inner electrons
(2) Closer packing among the nuclear particles is achieved
(3) The number of neutrons increases
(4) The increasing nuclear charge exerts a greater attractive force on the electrons

63. In an anion :-
(1) Number of proton decreases
(2) Protons are more than electrons
(3) Effective nuclear charge is more
(4) Radius is larger than neutral atom

64. Maximum size of first member of a period is due to
(1) Maximum number of shells
(2) Maximum screening effect
(3) Minimum Zeff
(4) All

65. Which of the following ion has largest size :-
(1) F^- (2) Al^{+3} (3) Cs^+ (4) O^{-2}

66. In which of the following pair radii of second species is smaller than that of first species :-
(1) Li, Na (2) Na^+, F^-
(3) N^{-3}, Al^{+3} (4) Mn^{+7}, Mn^{+4}

67. Spot the incorrect order of atomic radii :-
(1) $r_{Cu} > Zn$ (2) $r_{Cl} > F$ (3) $r_P > S$ (4) $r_{Sc} > Ti$

68. Which of the following orders of atomic radii are correct :-
(a) Li < Be < Na (b) Ni < Cu < Zn
(c) Ti > V > Cr (d) Ti > Zr ≃ Hf
Correct answer is :-
(1) All (2) a, b (3) b, c (4) b, d

69. Which electronic configuration of an atom is smallest in size :-
(1) $3s^2$ (2) $3s^2 3p^3$
(3) $3s^1$ (4) $3s^2\ 3p_x^2\ 3p_y^2\ 3p_z^1$

70. Decreasing order of size of ions is :-
(1) $Br^- > S^{-2} > Cl^- > N^{-3}$
(2) $N^{3-} > S^{-2} > Cl^- > Br^-$
(3) $Br^- > Cl^- > S^{-2} > N^{-3}$
(4) $N^{-3} > Cl^- > S^{-2} > Br^-$

71. Which of the following statement is wrong
(1) According to Slater, Z_{eff} in group remains constant
(2) In a period atomic size decreases
(3) Screening effect in a period remains constant
(4) In a period atomic radius of inert gas element is maximum

72. The covalent and vander Waal's radii of hydrogen respectively are :-
(1) 0.37 Å, 0.8 Å
(2) 0.37 Å, 0.37 Å
(3) 0.8 Å, 0.8 Å
(4) 0.8 Å, 0.37 Å

73. Which of the following sequence is correct for decreasing order of ionic radius :–
(1) Se^{-2}, I^-, Br^-, O^{-2}, F^-
(2) I^-, Se^{-2}, O^{-2}, Br^-, F^-
(3) Se^{-2}, I^-, Br^-, F^-, O^{-2}
(4) I^-, Se^{-2}, Br^-, O^{-2}, F^-

74. Element having maximum number of low shielding electrons :-
(1) [Xe] $4f^{14}$, $5d^{10}$, $6s^2$, $6p^2$
(2) [Rn] $5f^{14}$, $6d^1$, $7s^2$
(3) [Ar] $3d^{10}$, $4s^2$
(4) [Ne] $3s^2$, $3p^1$

75. Incorrect order of ionic radius is :-
(1) $La^{+3} > Gd^{+3} > Eu^{+3} > Lu^{+3}$
(2) $V^{+2} > V^{+3} > V^{+4} > V^{+5}$
(3) $In^{+} > Sn^{+2} > Sb^{+3}$
(4) $K^{+} > Sc^{+3} > V^{+5} > Mn^{+7}$

76. According to Slater's rule, order of effective nuclear charge for last electron in case of Li, Na and K :-
(1) Li > Na > K (2) K > Na > Li
(3) Na > Li > K (4) Li < Na = K

77. Rank the 4p, 4d and 4f orbitals of increasing order in which the electrons present in them are shielded by inner electrons
(1) 4d < 4f < 4p
(2) 4f < 4d < 4p
(3) 4p < 4d < 4f
(4) 4d < 4p < 4f

IONISATION POTENTIAL

78. Correct order of Ist I.P. are :-
(a) Li < B < Be < C
(b) O < N < F
(c) Be < N < Ne
(1) a, b (2) b, c (3) a, c (4) a, b, c

79. The ionisation potential of isotopes of an element will be :-
(1) Same
(2) Different
(3) Depends on atomic masses
(4) Depends on number of neutrons

80. The second ionisation potentials in electron volts of oxygen and fluorine atoms are respectively given by :-
(1) 35.1, 38.3
(2) 38.3, 38.3
(3) 38.3, 35.1
(4) 35.1, 35.1

81. A sudden large jump between the values of 2nd and 3rd IP of an element would be associated with the electronic configuration :-
(1) $1s^2, 2s^2\ 2p^6, 3s^1$
(2) $1s^2, 2s^2\ 2p^6, 3s^2\ 3p^5$
(3) $1s^2, 2s^2\ 2p^6, 3s^2\ 3p^2$
(4) $1s^2, 2s^2\ 2p^6\ 3s^2$

82. Compared to the first ionisation potential, the value of second ionisation potential of an element is :-
(1) Negligible (2) Smaller
(3) Greater (4) Double

83. In which of the following pairs, the ionisation energy of the first species is less than that of the second :-
(1) O^-, O^{2-} (2) S, P
(3) N, P (4) Be^+, Be

84. The correct order of stability of Al^+, Al^{+2}, Al^{+3} is :-
(1) $Al^{+3} > Al^{+2} > Al^+$ (2) $Al^{+2} > Al^{+3} > Al^+$
(3) $Al^{+2} < Al^+ > Al^{+3}$ (4) $Al^{+3} > Al^+ > Al^{+2}$

85. Least ionisation potential will be of :-
(1) Be^{3+} (2) H (3) Li^{+2} (4) He^+

86. Ionisation energy increases in the order :-
(1) Be, B, C, N (2) B, Be, C, N
(3) C, N, Be, B (4) N, C, Be, B

87. Mg forms Mg(II) because of :-
(1) The oxidation state of Mg is + 2
(2) Difference between $I.P_1$ and $I.P_2$ is greater than 16.0 eV
(3) There is only one electron in the outermost energy level of Mg
(4) Difference between $I.P_1$ and $I.P_2$ is less than 11 eV

88. Minimum first ionisation energy is shown by which electronic configuration:-
(1) $1s^2\ 2s^2\ 2p^5$
(2) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^2$
(3) $1s^2\ 2s^2\ 2p^6\ 3s^1$
(4) $1s^2\ 2s^2\ 2p^6$

89. With reference to ionisation potential which one of the following set is correct :-
(1) Li > K > B (2) B > Li > K
(3) Cs > Li > K (4) Cs < Li < K

90. Successive ionisation energies of an element 'X' are given below (in K. Cal)

IP_1	IP_2	IP_3	IP_4
165	195	556	595

Electronic configuration of the element 'X' is:-
(1) $1s^2, 2s^2 2p^6, 3s^2 3p^2$ (2) $1s^2, 2s^1$
(3) $1s^2, 2s^2 2p^2$ (4) $1s^2, 2s^2 2p^6, 3s^2$

91. Second IP of which of the element is maximum–
(1) Lithium (2) Oxygen
(3) Nitrogen (4) Fluorine

92. The energy needed to remove one electron from unipositive ion is abbreviated as :-
(1) 1st I.P. (2) 3rd I.P.
(3) 2nd I.P. (4) 1st E.A.

93. Among the following elements (Whose electronic configuration is given below) the one having the highest ionisation energy is
(1) $[Ne]\ 3s^2\ 3p^3$ (2) $[Ne]\ 3s^2\ 3p^4$
(3) $[Ne]\ 3s^2 3p^5$ (4) $[Ar]\ 3d^{10} 4s^2 4p^2$

94. The correct order of decreasing first ionisation energy is :-
(1) Si > Al > Mg > Na (2) Si > Mg > Al > Na
(3) Al > Si > Mg > Na (4) Mg > Li > Al > Si

95. Out of Na^+, Mg^{+2}, O^{-2} and N^{-3}, the pair of species showing minimum and maximum IP would be.
(1) Na^+, Mg^{+2} (2) Mg^{+2}, N^{-3}
(3) N^{-3}, Mg^{+2} (4) O^{-2}, N^{-3}

96. The element having highest I.P. is the from the two series C, N, O and Si, P, S :–
(1) P (2) N (3) S (4) O

97. Lowest IP will be shown by the element having the configuration :–
(1) [He] $2s^2$ (2) $1s^2$
(3) [He] $2s^2\ 2p^2$ (4) [He] $2s^2\ 2p^5$

98. The strongest reducing agent among the following is:-
(1) Na (2) Mg (3) Al (4) K

99. Which ionisation potential (IP) in the following equations involves the greatest amount of energy:-
(1) $K^+ \rightarrow K^{+2} + e^-$ (2) $Li^+ \rightarrow Li^{+2} + e^-$
(3) $Fe \rightarrow Fe^+ + e^-$ (4) $Ca^+ \rightarrow Ca^{+2} + e^-$

100. Values of first four ionisation potential of an elements are 68, 370, 400, 485. It belongs to which of the following electronic configuration:-
(1) $1s^2\ 2s^1$ (2) $1s^2\ 2s^2\ 2p^1$
(3) $1s^2\ 2s^2\ 2p^6\ 3s^1$ (4) (1) and (3) both

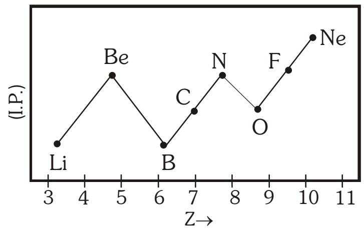
101. (a) $M^-_{(g)} \rightarrow M_{(g)}$ (b) $M_{(g)} \rightarrow M^+_{(g)}$
(c) $M^+_{(g)} \rightarrow M^{+2}_{(g)}$ (d) $M^{+2}_{(g)} \rightarrow M^{+3}_{(g)}$
Minimum and maximum I.P. would be of :-
(1) a, d (2) b, c (3) c, d (4) d, a

102. Which of the following electronic configuration belongs to least and most metallic character respectively:-
(a) $1s^2\ 2s^1$ (b) $5s^2\ 5p^5$
(c) $3s^2\ 3p^6\ 4s^1$ (d) $1s^2\ 2s^2\ 2p^5$
(1) a, b (2) d, c (3) b, a (4) c, d

103. Triad - I [N^{3-} , O^{-2}, Na^+]
Triad - II [N^+ , C^+ , O^+]
Choose the species of lowest IP from triad–I and highest IP from triad–II respectively
(1) N^{3-} , O^+ (2) Na^+ , C^+
(3) N^{3-} , N^+ (4) O^-, C^+

104. The correct values of ionisation energies (in kJ mol^{-1}) of Be, Ne, He and N respectively are
(1) 786, 1012, 999, 1256
(2) 1012, 786, 999, 1256
(3) 786, 1012, 1256, 999
(4) 786, 999, 1012, 1256

105. Following graph shows variation of I.P. with atomic number in second period (Li – Ne). Value of I.P. of Na (11) will be :-



(1) Above Ne
(2) Below Ne but above O
(3) Below Li
(4) Between N and O

106. Which one of the following has highest ionisation potential :–
(1) Li^+ (2) Mg^+ (3) He (4) Ne

107. In which of the following pairs, the ionisation energy of the first species is less than that of the second
(1) N, P (2) Be^+, Be
(3) N, N^- (4) Ne, Ne^+

108. Consider the following ionisation reactions
$A(g) \longrightarrow A^+(g) + e^-$ IE in (KJ/mol) is A_1
$A^+(g) \longrightarrow A^{+2}(g) + e^-$ IE in (KJ/mol) is A_2
$A^{+2}(g) \longrightarrow A^{+3}(g) + e^-$ IE in (KJ/mol) is A_3
then correct order of IE is :-
(1) $A_1 > A_2 > A_3$ (2) $A_1 = A_2 = A_3$
(3) $A_1 < A_2 < A_3$ (4) $A_3 = A_2 < A_1$

109. IE_1 IE_2 and IE_3 of an element are 10 eV, 15 eV, 45 eV respectively, the most stable oxidation state of the element will be :-
(1) + 1 (2) + 2 (3) + 3 (4) + 4

110. Select the correct order of I.E. :-
(1) $Cl^- > Cl > Cl^+$ (2) $Cl^+ > Cl > Cl^-$
(3) $Cl > Cl^+ > Cl^-$ (4) $Cl^- > Cl^+ > Cl$

ELECTRON AFFINITY

111. In the process $Cl_{(g)} + e^- \xrightarrow{\Delta H} Cl^-(g)$, ΔH is
(1) Positive (2) Negative
(3) Zero (4) None

112. Process in which maximum energy is released:-
(1) $O \rightarrow O^{-2}$ (2) $Mg^+ \rightarrow Mg^{+2}$
(3) $Cl \rightarrow Cl^-$ (4) $F \rightarrow F^-$

113. Which of the following is energy releasing process
(1) $X^- \rightarrow X\ (g) + e^-$
(2) $O^-(g) + e^- \rightarrow O^{2-}$
(3) $O\ (g) \rightarrow O^+\ (g) + e^-$
(4) $O\ (g) + e^- \rightarrow O^-(g)$

114. In which of the following process energy is liberated:-

(1) $Cl \rightarrow Cl^+ + e$ (2) $HCl \rightarrow H^+ + Cl^-$

(3) $Cl + e \rightarrow Cl^-$ (4) $O^- + e \rightarrow O^{-2}$

115. Element of which atomic number has highest electron affinity:-

(1) 35 (2) 17 (3) 9 (4) 53

116. The electron affinity

(1) Of carbon is greater than oxygen

(2) Of fluorine is less than iodine

(3) Of fluorine is less than chlorine

(4) Of sulphur is less than oxygen

117. Which of the following element will form most stable bivalent anion.

(1) Fluorine (2) Oxygen

(3) Chlorine (4) Nitrogen

118. Energy absorbed in second electron addition in an atom is called.

(1) 1^{st}IP (2) 2^{nd}EA

(3) 1^{st}EA (4) 2^{nd}IP

119. The amount of energy released for the process $X_{(g)} + e^- \rightarrow X^-_{(g)}$ is minimum and maximum respectively for :–

(a) F (b) Cl (c) N (d) B

Correct answer is :–

(1) c & a (2) d & b

(3) a & b (4) c & b

120. Which of the following electronic configuration is expected to have highest electron affinity:-

(1) $2s^2\ 2p^0$ (2) $2s^2\ 2p^2$

(3) $2s^2\ 2p^3$ (4) $2s^2\ 2p^1$

121. Consider the following conversions

(i) $O(g) + e^- \longrightarrow O^-(g)\ ;\ \Delta H_1$

(ii) $F(g) + e^- \longrightarrow F^-(g)\ ;\ \Delta H_2$

(iii) $Cl(g) + e^- \longrightarrow Cl^-(g)\ ;\ \Delta H_3$

(iv) $Na(g) \longrightarrow Na^+(g)\ ;\ \Delta H_4$

incorrect statement is :-

(1) ΔH_1 and ΔH_2 is less negative than ΔH_3

(2) ΔH_2 is more negative than ΔH_1

(3) ΔH_2 , ΔH_3 are negative while ΔH_1 is positive

(4) ΔH_1 , ΔH_2 and ΔH_3 are negative while ΔH_4 is positive

122. In which of the following process, least energy is required :-

(1) $F^-_{(g)} \longrightarrow F_{(g)} + e^-$

(2) $P^-_{(g)} \longrightarrow P_{(g)} + e^-$

(3) $S^-_{(g)} \longrightarrow S_{(g)} + e^-$

(4) $Cl^-_{(g)} \longrightarrow Cl_{(g)} + e^-$

ELECTRONEGATIVITY

123. The correct set of decreasing order of electronegativity is :-

(1) Li, H, Na (2) Na, H, Li

(3) H, Li, Na (4) Li, Na, H

124. Polarity of a bond can be explained by :–

(1) Electron affinity (2) Ionisation potential

(3) Electronegativity (4) All of the above

125. Electronegativity values for elements are useful in predicting :-

(1) Bond energy of a molecule

(2) Polarity of a bond

(3) Nature of an oxide

(4) All

126. Mulliken scale of electronegativity uses the concept of :-

(1) E. A. and EN of pauling

(2) E. A. and atomic size

(3) E.A. and I.P.

(4) E.A. and bond energy

127. The pair with minimum difference in electronegativity is :-

(1) F, Cl (2) C,H (3) P, H (4) Na, Cs

128. Least electronegative element is :-

(1) I (2) Br (3) C (4) Cs

129. In which of the following pairs of elements the electronegativity of first element is less than that of second element :-

(1) Zr, Hf

(2) K, Rb

(3) Cl, S

(4) None of the above

130. The nomenclature of ICl is iodine chloride because

(1) Size of I < Size of Cl

(2) Atomic number of I > Atomic number of Cl

(3) E.N. of I < E.N. of Cl

(4) E. A. of I < E. A. of Cl

131. Among the following least and most polar bonds are respectively :-

(a) C – I (b) N – O (c) C – F (d) P – F

(1) d and c (2) a and d (3) b and d (4) b and c

132. If the ionisation potential is IP, electron affinity is EA and electronegativity is X then which of the following relation is correct :-

(1) 2X – EA – IP = 0

(2) 2EA – X – IP = 0

(3) 2IP – X – EA = 0

(4) All of the above

133. The properties which are not common to both groups 1 and 17 elements in the periodic table are :-
(1) Electropositive character increases down the groups
(2) Reactivity decreases from top to bottom in these groups
(3) Atomic radii increases as the atomic number increases
(4) Electronegativity decreases on moving down the group

134. Electronegativity of an element can be measured using :-
(1) Pauling's scale (2) Mulliken's scale
(3) Both (4) None

135. As we proceed across the period in periodic table, we find there is a decrease in :-
(1) Ionisation energy (2) Electron affinity
(3) Electronegativity (4) Atomic radii

136. Which compound strongly absorb CO_2 ?
(1) BeO (2) K_2O
(3) H_3PO_4 (4) P_4O_6

137. The electronegativities of the following elements: H, O, F, S and Cl increase in the order :-
(1) H < O < F < S < Cl
(2) Cl < H < O < F < S
(3) H < S < O < Cl < F
(4) H < S < Cl < O < F

138. Which of the following is different from other three oxides :-
(1) MgO (2) SnO
(3) PbO (4) ZnO

EXERCISE-I ANSWER KEY

Que.	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15
Ans.	**4**	**3**	**1**	**1**	**2**	**2**	**4**	**2**	**4**	**3**	**2**	**1**	**3**	**2**	**2**
Que.	16	17	18	19	20	21	22	23	24	25	26	27	28	29	30
Ans.	**2**	**1**	**2**	**2**	**3**	**1**	**1**	**1**	**3**	**4**	**1**	**3**	**1**	**1**	**2**
Que.	31	32	33	34	35	36	37	38	39	40	41	42	43	44	45
Ans.	**3**	**3**	**3**	**2**	**1**	**2**	**4**	**1**	**3**	**2**	**2**	**3**	**3**	**2**	**1**
Que.	46	47	48	49	50	51	52	53	54	55	56	57	58	59	60
Ans.	**1**	**3**	**3**	**2**	**1**	**2**	**3**	**4**	**3**	**1**	**1**	**2**	**2**	**4**	**1**
Que.	61	62	63	64	65	66	67	68	69	70	71	72	73	74	75
Ans.	**4**	**4**	**4**	**3**	**3**	**3**	**1**	**3**	**4**	**1**	**3**	**1**	**4**	**2**	**1**
Que.	76	77	78	79	80	81	82	83	84	85	86	87	88	89	90
Ans.	**4**	**3**	**4**	**1**	**3**	**4**	**3**	**2**	**4**	**2**	**2**	**4**	**3**	**2**	**4**
Que.	91	92	93	94	95	96	97	98	99	100	101	102	103	104	105
Ans.	**1**	**3**	**3**	**2**	**3**	**2**	**1**	**4**	**2**	**3**	**1**	**2**	**1**	**3**	**3**
Que.	106	107	108	109	110	111	112	113	114	115	116	117	118	119	120
Ans.	**1**	**4**	**3**	**2**	**2**	**2**	**3**	**4**	**3**	**2**	**3**	**2**	**2**	**4**	**2**
Que.	121	122	123	124	125	126	127	128	129	130	131	132	133	134	135
Ans.	**3**	**2**	**3**	**3**	**4**	**3**	**3**	**4**	**1**	**3**	**2**	**1**	**2**	**3**	**4**
Que.	136	137	138												
Ans.	**2**	**4**	**1**												

Directions for Assertion & Reason questions

These questions consist of two statements each, printed as Assertion and Reason. While answering these Questions you are required to choose any one of the following four responses.

(A) If both Assertion & Reason are True & the Reason is a correct explanation of the Assertion.

(B) If both Assertion & Reason are True but Reason is not a correct explanation of the Assertion.

(C) If Assertion is True but the Reason is False.

(D) If both Assertion & Reason are false.

1. ***Assertion*** : Two successive ionisation energies of Argon are 56.8 eV and 36.8 eV respectively.
Reason : Zeff of Ar ($3s^2 3p^6$) is greater than Ar^+ ($3s^2 3p^5$).
(1) A (2) B (3) C (4) D

2. ***Assertion*** : Electron affinity of fluorine is greater than chlorine.
Reason : Ionisation potential of fluorine is less than chlorine.
(1) A (2) B (3) C (4) D

3. ***Assertion*** : IP of F > EA of F
Reason : F is a highly electronegative element
(1) A (2) B (3) C (4) D

4. ***Assertion*** : Properties of Beryllium is similar to that of Aluminium
Reason : Both the elements belongs to same group
(1) A (2) B (3) C (4) D

5. ***Assertion*** : I.P. of first element in a period is minimum.
Reason : Effective nuclear charge of first element in a period is minimum
(1) A (2) B (3) C (4) D

6. ***Assertion*** : Size of anion is larger than its parent atom.
Reason : Z_{eff} of anion is greater than that of their parent atom.
(1) A (2) B (3) C (4) D

7. ***Assertion*** : Stable electronic configuration does not affects electronegativity.
Reason : EN is the tendency of an atom to attract shared electrons, not to gain electrons.
(1) A (2) B (3) C (4) D

8. ***Assertion*** : Chlorine is most electronegative element.
Reason : Chlorine has tendency to loose electrons.
(1) A (2) B (3) C (4) D

9. ***Assertion*** : Atomic radius increases, descending down the group.
Reason : On going down the group EN increase.
(1) A (2) B (3) C (4) D

10. ***Assertion*** : Atomic radius of inert gases are largest in the period
Reason : Effective nuclear charge of inert gases are minimum
(1) A (2) B (3) C (4) D

11. ***Assertion*** : Second IP of oxygen is greater than that of fluorine
Reason : Oxygen aquires stable half filled electronic configuration after loosing one electron
(1) A (2) B (3) C (4) D

12. ***Assertion*** : Electronegativity of nitrogen is greater than carbon.
Reason : Nitrogen has stable half filled electronic configuration.
(1) A (2) B (3) C (4) D

13. ***Assertion*** : Atomic size of Boron is larger than Beryllium
Reason : Number of shell in Boron is greater than Beryllium
(1) A (2) B (3) C (4) D

14. ***Assertion*** : Alkali metals have least 1^{st} I.P. in the respective period
Reason : Alkali metals have only one electron in the valence shell.
(1) A (2) B (3) C (4) D

15. ***Assertion*** : Atomic size of Na is larger than Mg.
Reason : No. of shell in Mg is more than Na.
(1) A (2) B (3) C (4) D

16. ***Assertion*** : Na^+ and Cl^- have similar ionic radius.
Reason : Z_{eff} in Na^+ and Cl^- are same.
(1) A (2) B (3) C (4) D

17. ***Assertion :*** Ionisation potential of Li^+ is greater than He.

Reason : Zeff of Li^+ is greater than He.

(1) A (2) B (3) C (4) D

18. ***Assertion :*** Size of Ca^{+2} is larger than K^+.

Reason : Number of electrons in Ca^{+2} is more than K^+.

(1) A (2) B (3) C (4) D

19. ***Assertion :*** 2^{nd} IP of alkali metals is maximum in the period.

Reason : Alkali metals have the smallest atomic size in the period.

(1) A (2) B (3) C (4) D

20. ***Assertion :*** Atomic size along a period decreases.

Reason : Zeff. in a period decreases.

(1) A (2) B (3) C (4) D

21. ***Assertion :*** The 1^{st} IP of Be is greater than that of B.

Reason : 2p orbital is lower in energy than 2s.

(1) A (2) B (3) C (4) D

22. ***Assertion :*** First ionization energy of nitrogen is lower than oxygen. **[AIIMS-2005]**

Reason : Across the period effective nuclear charge decreases.

(1) A (2) B (3) C (4) D

23. **Assertion :-** H_2Se is less acidic than H_2S.

Reason :- S is less electronegative than Se.

(1) A (2) B (3) C (4) D

24. ***Assertion :*** Atomic sizes of Cs and Fr are almost similar.

Reason : Cs and Fr belong to same group.

(1) A (2) B (3) C (4) D

25. ***Assertion :-*** Zn is not a transition element whereas Sc is

Reason :- Outershell configuration of Zn is $3d^{10}$, $4s^2$, $4p^0$, $4d^0$.

(1) A (2) B (3) C (4) D

EXERCISE-II ANSWER KEY

Que.	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15
Ans.	**4**	**4**	**2**	**3**	**1**	**3**	**1**	**4**	**3**	**3**	**1**	**2**	**4**	**2**	**3**
Que.	16	17	18	19	20	21	22	23	24	25					
Ans.	**4**	**1**	**4**	**3**	**3**	**3**	**4**	**4**	**2**	**2**					